

Chapter \rightarrow States of Matter

Intermolecular forces

\rightarrow Attractive forces b/w dipoles of molecules or species are called Van der Waal forces.

\rightarrow D-D force (Dipole-Dipole interaction)
(Keom forces)

\rightarrow D-ID forces (Dipole-Induced)
(Debye forces) (dipole forces)

\rightarrow Instantaneous Dipole-Instantaneous Dipole forces
(ID-ID forces) } Weakest force of nature

or
London
Dispersive forces

Dipole \rightarrow Used for Polar molecule ($\mu \neq 0$)

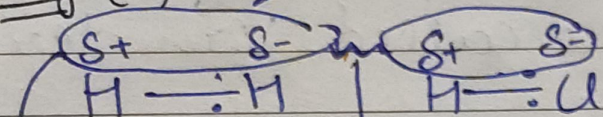
\rightarrow Represent Permanent Dipole

egⁿ CHCl_3 , H_2O , NH_3 , NF_3 , BrF_5 , HCl etc

Temporary Dipole \rightarrow either induced or Instantaneous

Dipole is created in any non-polar molecule or noble gas atom by some polar molecule present in surrounding

egⁿ (i) $H_2 + HCl$

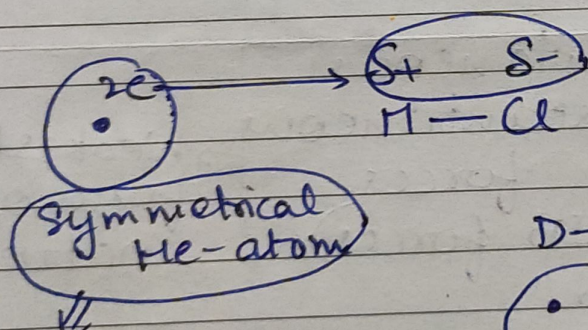


induced due to dipole of HCl

D-ID force

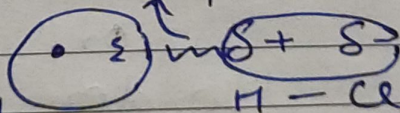
created b/w 2 non-polar species or Noble gas atoms due to disturbance of e^- density.

(ii) $He + HCl$

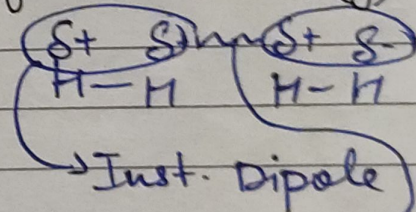


loses its symmetry due to disturbance caused by dipole of HCl

D-ID force



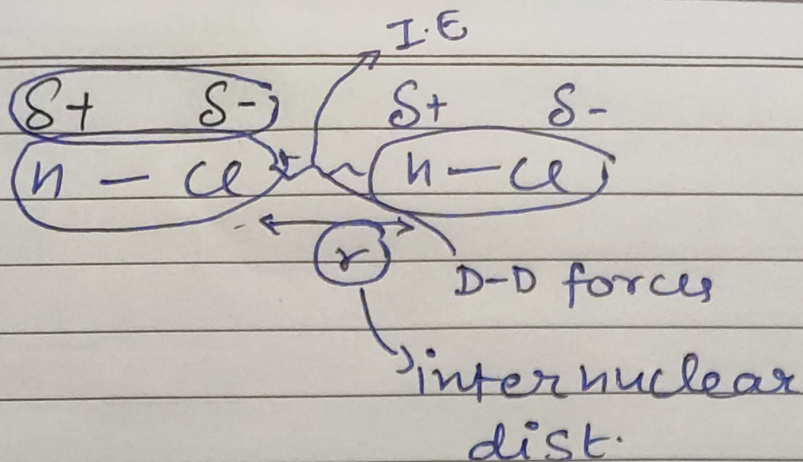
egⁿ forces in $H_2(g)$



Inst. Dipole
ID-ID forces or London Dispersive forces

Interaction Energy released (I.E)

\propto Strength of Dipole & their Interaction

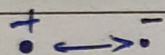
egⁿ

$$|I \cdot E| \propto \frac{1}{r^3} \quad \left(\text{Dipole are stationary} \right)$$

Solid

$$|I \cdot E| \propto \frac{1}{r^6} \quad \left(\text{,, ,, moving, rotating} \right)$$

Liq, gas



$$\left| \frac{k q_1 q_2}{r^2} \right|$$

$$D-ID \Rightarrow |I \cdot E| \propto \frac{1}{r_1^6}$$

$$ID-ID \rightarrow |I \cdot E| \propto \frac{1}{r_2^6}$$

$$r_3 < r_1 < r_2$$

$$(D-D)r < (D-ID)r < (ID-ID)r$$

I. E order

$$D-D > D-ID > ID-ID$$

Chapter → States of Matter

★ Common State of Matter

① Solid state

② Liq. state

③ Gaseous state

→ Factors to differentiate

→ IMF

① Intermolecular forms (Attraction)

↳ Bring Particles closer

② Thermal Energy (T.E)

↳ ↑ T.E of particles so particles move away from each other.

• if $TE \gg IMF$ → Gaseous state

• if $T.E > IMF$ → Liq. state

• if $IMF \gg TE$ → Solid state

$1 \text{ dm}^3 = 1 \text{ Litre}$

(#) Gaseous state of Matter

To discuss state (or conditions) of any gas

Measurable parameters for any Gas will be

1) Pressure 2) Volume 3) Temperature

4) Amount of Gas

↳ moles, mass of Gas

→ & considering Every gas to be perfect or Ideal, we have ideal Gas eqn or Eqn of state of Ideal Gas

$$\boxed{P \times V = n \times R \times T}$$

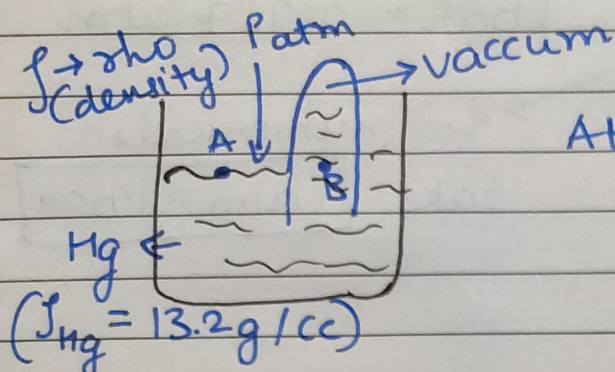
\downarrow Pressure of Gas \downarrow vol. of Gas \downarrow moles of Gas \downarrow Universal Gas constant \rightarrow Temp. of Gas

(**) Pressure of Gas \rightarrow only due to collision of Gas molecules on the wall of container or vessel.

↓
More is total force acting per unit area of wall is called pressure of Gas.

Atmospheric pressure

↳ is measured by Barometer.



At same level of liq,
same Pressure is observed.

$$P_A = P_B$$

$$P_{atm} = \rho_{liq} \cdot g \cdot h_{liq}$$

$$P_A = P_{atm}$$

$$P_B = \rho \cdot g \cdot h$$

$$P_B = \rho_{Hg} \cdot g \cdot h_{Hg}$$

force acting
on Unit
area

When $P_{atm} = 1 \text{ atm}$

then Height of Hg = 76 cm Hg

& Height of Hg is taken as unit of Pressure.

$$1 \text{ atm} = 76 \text{ cm Hg}$$

$$1 \text{ atm} = 760 \text{ mm Hg}$$

$$1 \text{ atm} = 760 \text{ torr}$$

$$1 \text{ mm Hg} = 1 \text{ torr}$$

⊕

$$1 \text{ Pascal (Pa)} = \frac{1 \text{ N}}{\text{m}^2}$$

$$1 \text{ bar} = \frac{10^5 \text{ N}}{\text{m}^2} \quad \text{or} \quad 10^5 \text{ Pa}$$

$$1 \text{ bar} = 100 \text{ kPa}$$

$$1 \text{ kPa} = 1 \text{ Kilopascal} = 10^3 \text{ Pa}$$

$$1 \text{ atm} > 1 \text{ bar}$$

$$1 \text{ atm} = 1.01325 \text{ bar}$$

$$1 \text{ bar} = 0.987 \text{ atm}$$

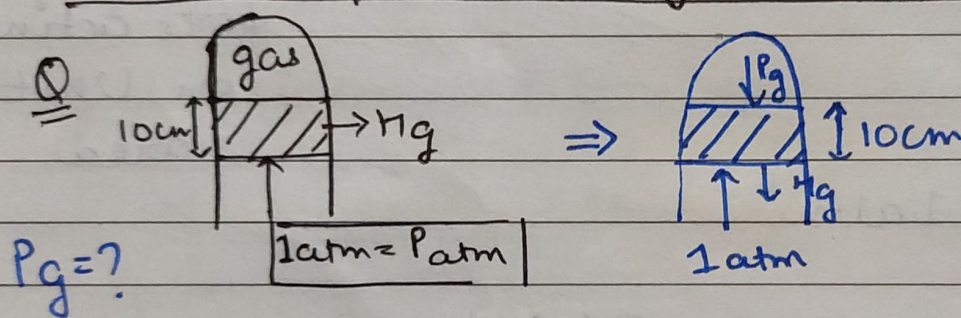
$$1 \text{ atm} = 1.01325 \times 10^5 \text{ Pa}$$

$$\approx 1.01325 \times 100 \text{ kPa}$$

for numericals
take $1 \text{ atm} = 1 \text{ bar}$

$$1 \text{ atm} = 101.325 \text{ kPa}$$

Pressure of Gas using Faulty Barometer



$$P_{\text{atm}} = P_g + P_{\text{Hg}}$$

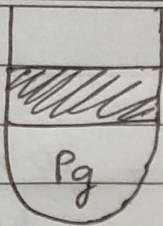
$$76 = \cancel{8} + P_g + 10$$

$$66 = P_g$$

$$66 \text{ cm Hg} = P_g$$

$$P_g = \frac{66}{76} \text{ atm}$$

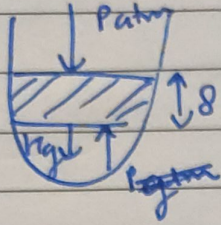
Q →



8 cm Hg

$$P_{atm} = 1 \text{ atm}$$

$$P_g = ?$$

Solⁿ

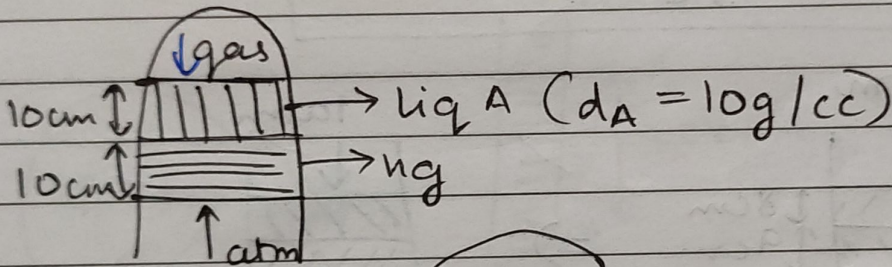
$$P_{atm} + P_{hg} = P_g$$

$$76 + 8 = P_g$$

$$84 \text{ cm Hg} = P_g$$

$$P_g = \frac{84}{76} \text{ atm}$$

Q →



$$P_g = ?$$

Solⁿ

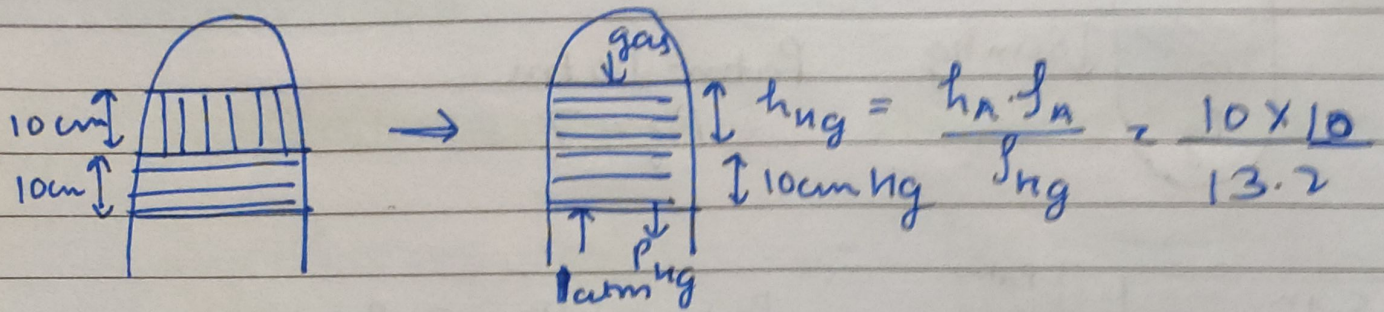
Convert other liquid height into height of Hg for same equivalent Pressure.

$$P_{liq} = P_{Hg}$$

$$\rho_{liq} \cdot g \cdot h_{liq} = \rho_{Hg} \cdot g \cdot h_{Hg}$$

$$h_{Hg} =$$

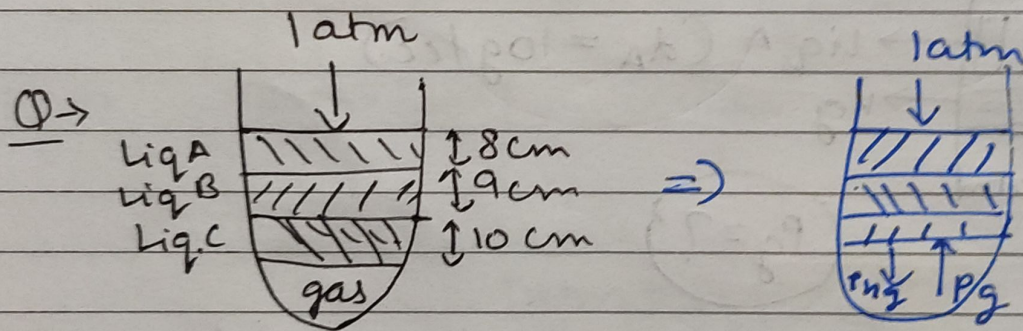
$$\frac{\rho_{liq} \cdot h_{liq}}{\rho_{Hg}}$$



$$76 = P_g + P_{hg}$$

$$76 - P_{hg} = P_g$$

$$P_g = \left(76 - \frac{100}{13.2} - 10 \right) \text{ cm hg}$$



$$P_g = ?$$

$$\begin{aligned} \rho_{hg} &= 13.2 \text{ g/cc} \\ \rho_A &= 6 \text{ g/cc} \\ \rho_B &= 8 \text{ g/cc} \\ \rho_C &= 4 \text{ g/cc} \end{aligned}$$

$$h_{hg} = \frac{\rho_A \cdot h_A}{\rho_{hg}} = \frac{6 \cdot 8}{13.2}$$

$$h_{hg} = \frac{\rho_B \cdot h_B}{\rho_{hg}} = \frac{8 \cdot 9}{13.2}$$

$$h_{hg} = \frac{\rho_C \cdot h_C}{\rho_{hg}} = \frac{4 \cdot 10}{13.2}$$

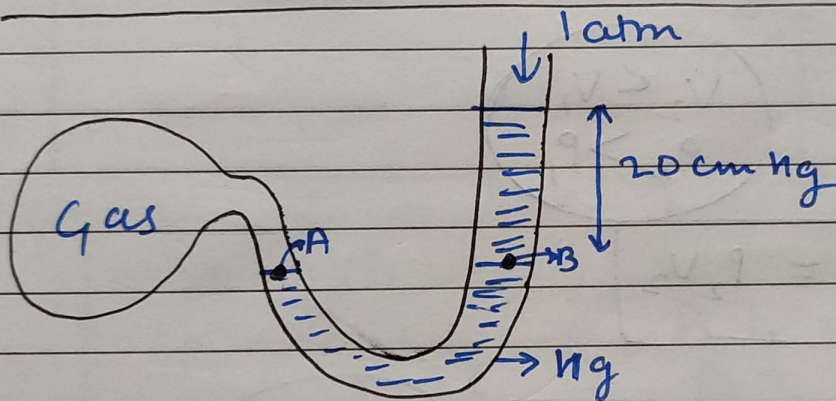
$$P_g = P_{atm} + P_{Hg}$$

$$= 76 + \frac{48}{13.2} + \frac{72}{13.2} + \frac{40}{13.2}$$

$$= \frac{1003.2 + 48 + 72 + 40}{13.2}$$

$$= \frac{1163.2}{13.2} \Rightarrow 88 \text{ (Approx)}$$

(*) Another Device to measure P_g is Manometer



$$P_A = P_B$$

$$P_g = 1 \text{ atm} + P_{Hg}$$

$$= 76 + 20$$

$$= 96 \text{ cm Hg}$$

$$P_g = \frac{96}{76} \text{ atm} \text{ or } \frac{96}{76} \times 101.325 \text{ kPa}$$

(#) Ideal Gas Law

$$P V = n R T$$

i) Boyle's law \div at const. T \rightarrow Isothermal State
& at const. n

Date

$$P = \frac{nRT}{V}$$

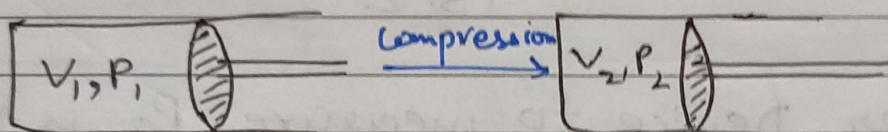
const. (k)

Vol. of ideal Gas
= vol. of Container
in which Gas
is present

$P \uparrow, V \downarrow$, compression

$$P \propto \frac{1}{V}$$

$P \downarrow, V \uparrow$ Expansion



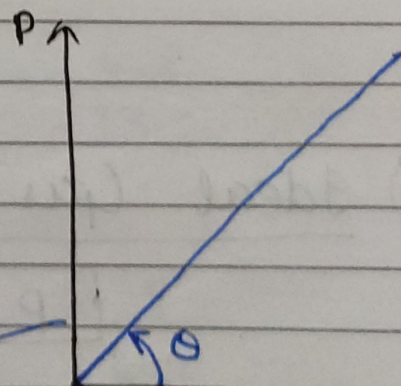
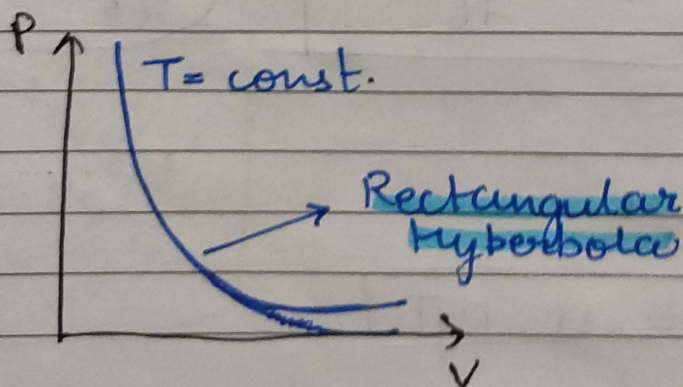
$$V_2 < V_1$$

$$P_2 > P_1$$

$$P_1 V_1 = P_2 V_2$$

$$PV = \text{const. (k)}$$

$$P \propto \frac{1}{V}$$

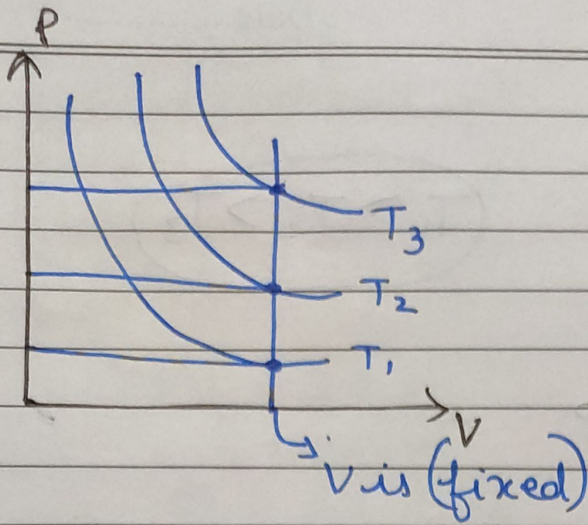


$$\text{slope} = \tan \theta = k = nRT$$

$$P = \frac{nRT}{V}$$

$$\begin{matrix} (P) & (k) & (1) \\ y & \downarrow m & (V) \rightarrow x \end{matrix}$$

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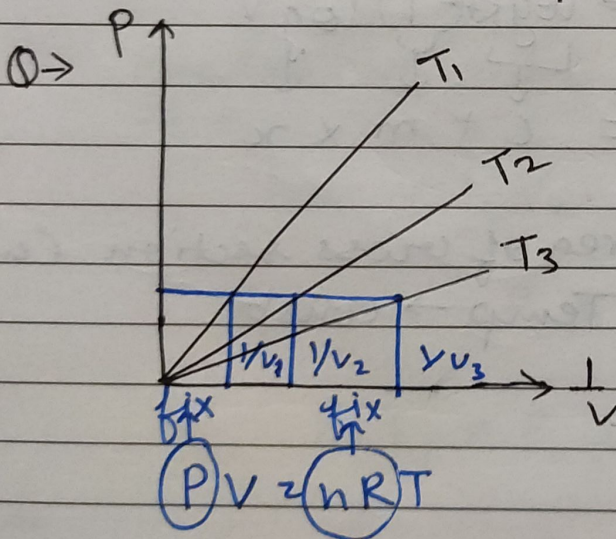
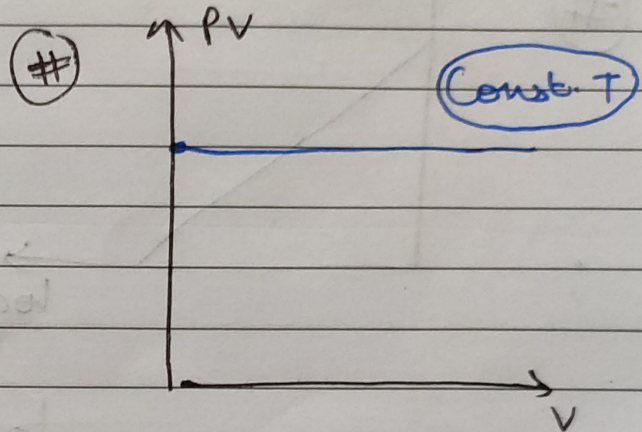
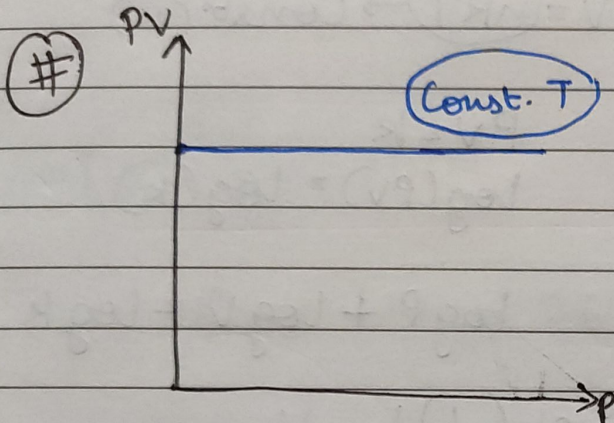
$$PV = nRT$$

\downarrow fix \downarrow fix \downarrow fix

$$P \propto T$$

$$P_3 > P_2 > P_1$$

$$T_3 > T_2 > T_1$$



$$V = \left(\frac{nR}{P} \right) T$$

$$V \propto T$$

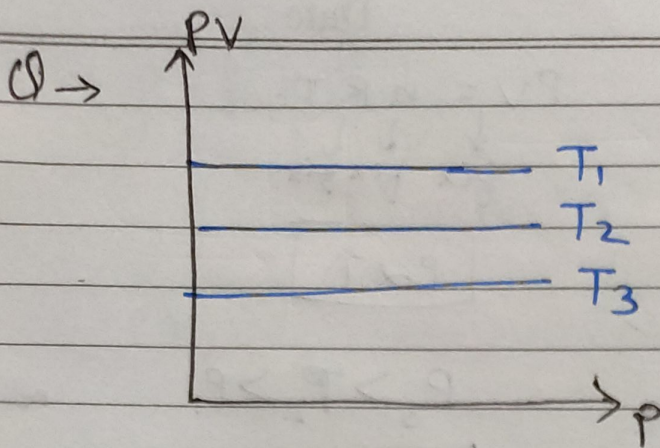
$$\frac{1}{V} \propto \frac{1}{T}$$

- 1) $T_1 > T_2 > T_3$
- 2) $T_2 > T_3 > T_1$
- 3) $T_3 > T_2 > T_1$
- 4) $T_3 > T_1 > T_2$

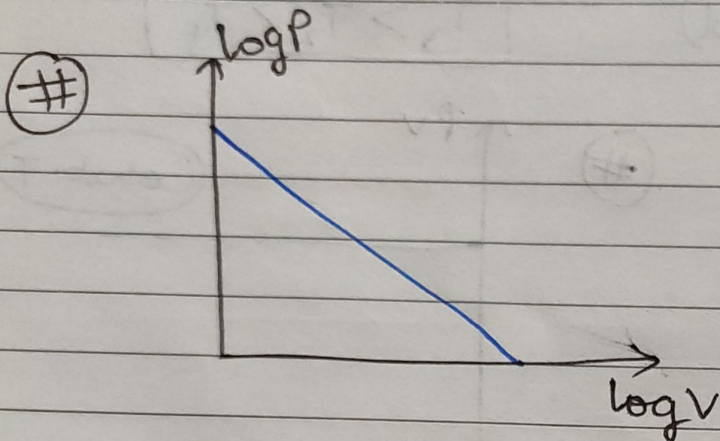
$$\frac{1}{V_3} > \frac{1}{V_2} > \frac{1}{V_1}$$

$$\frac{1}{T_3} > \frac{1}{T_2} > \frac{1}{T_1}$$

$$T_1 > T_2 > T_3$$



$$T_1 > T_2 > T_3$$



$$PV = nRT \rightarrow \text{const. (K)}$$

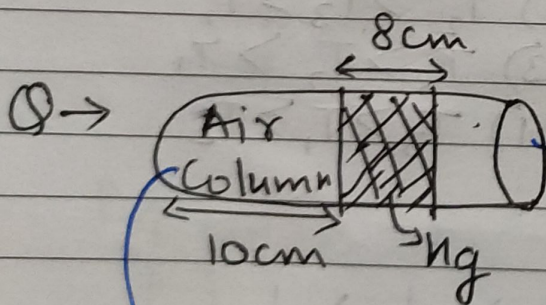
$$PV = K$$

$$\log(PV) = \log(K)$$

$$\log P + \log V = \log K$$

$$\underbrace{\log P}_{y} = \underbrace{\log K}_{c} + \underbrace{(-1)}_{m} \underbrace{\log V}_{x}$$

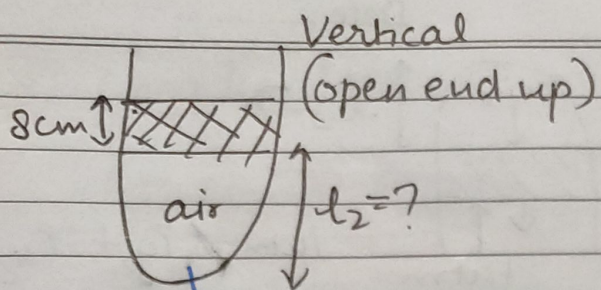
$$y = c + m \times x$$



$A = \text{Area of cross section (cm}^2\text{)}$
 Temp \rightarrow const

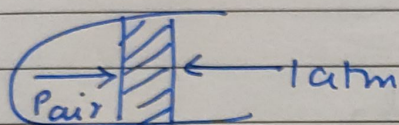
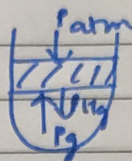
$$V_1 = A \times l_1 = (A \times 10)$$

$$P_{air} = P_1 = 76 \text{ cm Hg}$$



$$V_2 = A \times l_2$$

$$P_2 = 84 \text{ cmHg}$$



$$P_{\text{air}} = 1 \text{ atm} = 76 \text{ cmHg}$$

Boyle's law for air can be applied

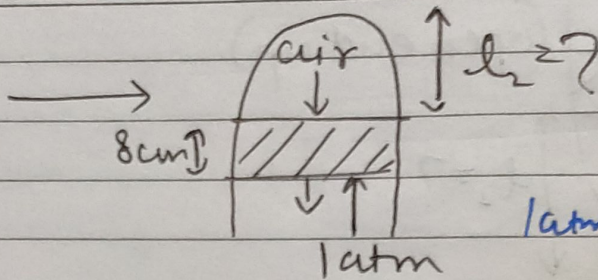
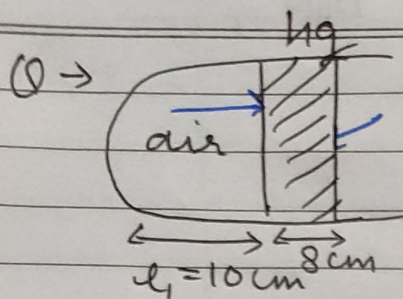
$$P_1 V_1 = P_2 V_2$$

$$P_1 A l_1 = P_2 A l_2$$

$$P_1 l_1 = P_2 l_2$$

$$l_2 = \frac{P_1 l_1}{P_2} = \left(\frac{76 \times 10}{84} \right) \text{ cm}$$

$$= \frac{760}{84} \approx 9 \text{ (Approx)}$$



$$1 \text{ atm} = P_g + P_{Hg}$$

$$P_{\text{atm}} = P_g$$

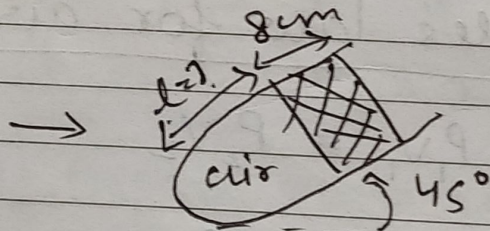
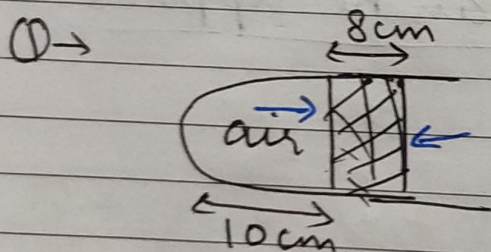
$$76 - 8 = P_g$$

$$P_g = 68 \text{ cmHg}$$

$$P_g = 76 \text{ cmHg}$$

$$P_1 \times l_1 = P_2 \times l_2$$

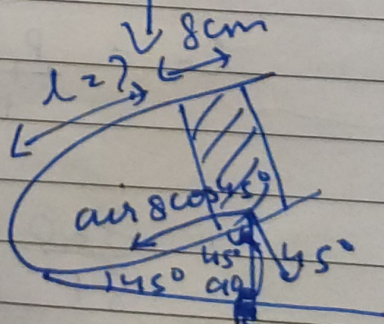
$$\text{cm} \left(\frac{10 \times 76}{68} \right) = l_2$$



$$P_{\text{atm}} = P_{\text{air}} = 76 \text{ cmHg}$$

$$P_1 l_1 = P_2 l_2$$

$$\left(\frac{76 \times 10}{76 + 4\sqrt{2}} \right) \text{ cm} = l_2$$



$$P_g = 76 + 8 \cos 45^\circ$$

$$P_g = 76 + 4\sqrt{2}$$

② Charles Law :-

At const. P, n

Volume of Gas $\propto T$

Const. $P \rightarrow$ Isobaric
condⁿ

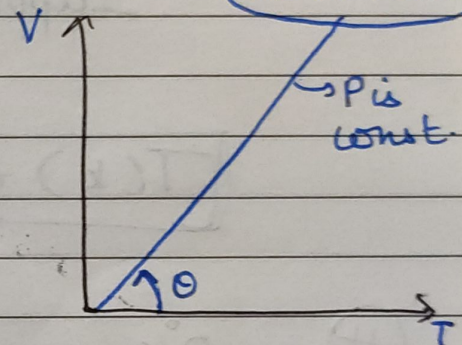
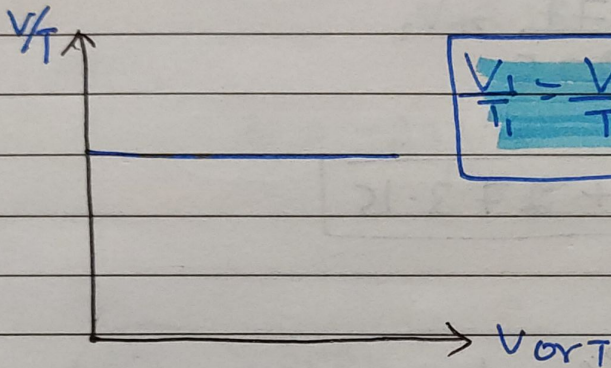
$$V = \left(\frac{nR}{P} \right) T$$

\hookrightarrow const. (k)

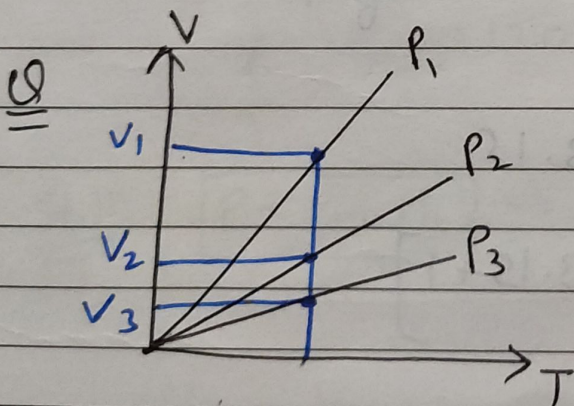
$$\frac{V}{T} = \text{const. } (k)$$

$$V = kT$$

T is in kelvin



$$\text{slope} = \tan \theta = k = \frac{nR}{P}$$



$$V \propto \frac{1}{P}$$

$$P_1 < P_2 < P_3$$

$$V_1 > V_2 > V_3$$

Q \Rightarrow find Volume of Gas? if Temp. of 1 litre Gas is increased from 2°C to 4°C keeping P constant.

Solⁿ :- $t_1 = 2^\circ\text{C}$
 $V_1 = 1 \text{ litre}$
 Initial condⁿ $\xrightarrow{P = \text{const.}}$ $\boxed{4^\circ\text{C} = t_2}$
 $V_2 = ?$ \rightarrow final condⁿ

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{1}{2+273} = \frac{V_2}{4+273}$$

$$1 (\text{approx}) = \frac{277}{275} = V_2$$

$$\boxed{T(K) = t^\circ\text{C} + 273.15}$$

(#) $0^\circ\text{C} \longrightarrow 1^\circ\text{C} = t$
 $\swarrow V_0$ $V_t = V_1$
 Vol. of Gas at 0°C \hookrightarrow Vol. of Gas at 1°C

$$\frac{V_0}{0+273.15} = \frac{V_1}{1+273.15}$$

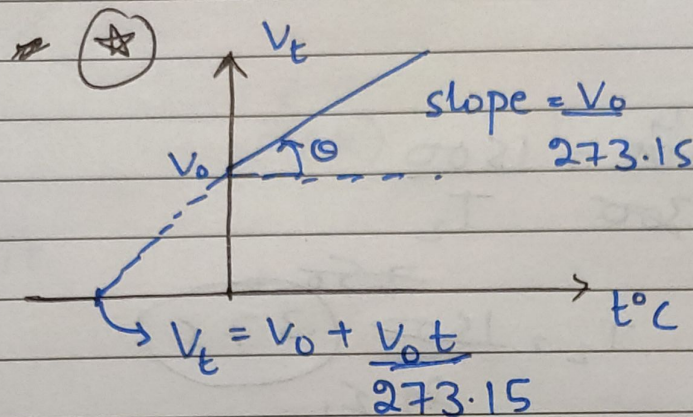
$$V_1 = \frac{V_0}{273.15} [273.15 + 1]$$

$$V_1 = V_0 + \frac{V_0}{273.15}$$

T in Kelvin
 $[\Delta t \text{ or } \Delta T] \rightarrow$
 or
 $^\circ\text{C}$

$$0^{\circ}\text{C} \longrightarrow t^{\circ}\text{C}$$

$$V_t = V_0 + \left(V_0 \times \frac{(t-0)}{273.15} \right) = V_0 + \frac{V_0 \Delta t}{273.15}$$

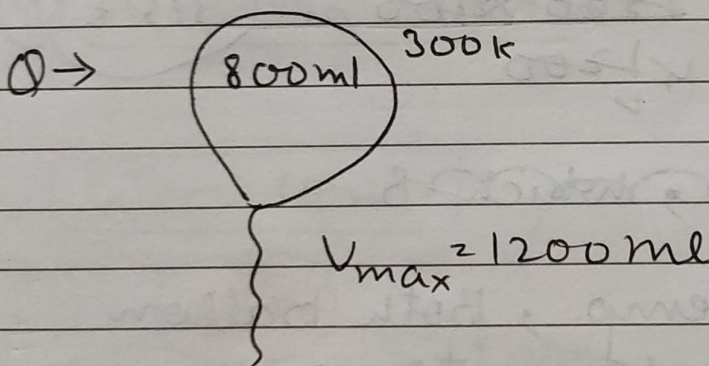


$$V_t = V_0 + V_0 \frac{t}{273.15}$$

\downarrow y \downarrow C \downarrow m

$$t = -273.15$$

\hookrightarrow 0 K (Absolute zero)



Ballon is taken to a room ^{where} temp. is \uparrow ing. At what temp. Ballon will Burst?

Solⁿ $P = \text{const.}$ \Rightarrow during Bursting of Ballon or tyre

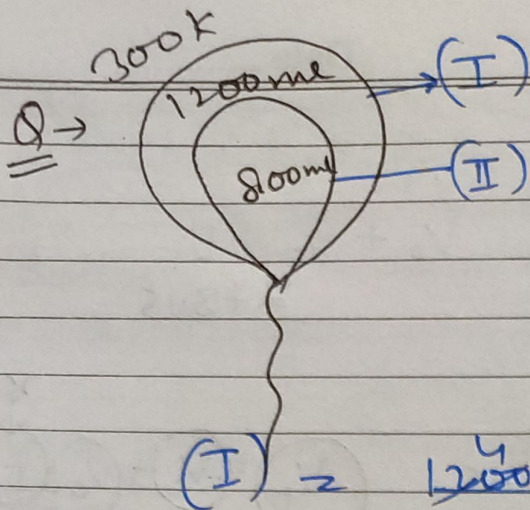
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Ballon Burst

$$T > 450\text{K}$$

$$\frac{800}{300} = \frac{1200}{T_2}$$

$$T_2 = \frac{1200 \times 300}{800} = 450\text{K}$$



$$V_{\max}(I) = 1500 \text{ ml}$$

$$V_{\max}(II) = 1000 \text{ ml}$$

$$(I) = \frac{1200}{300} = \frac{1500}{T_2}$$

$$T_2 = \frac{1500}{1200} \times 300 = 375$$

$$T(II) = 375 \text{ (K)}$$

$$T(I) = \frac{1500}{1200} \times 300 = 375 \text{ K}$$

on same Temp, which is

On same Temp., Both balloon burst on same temp.

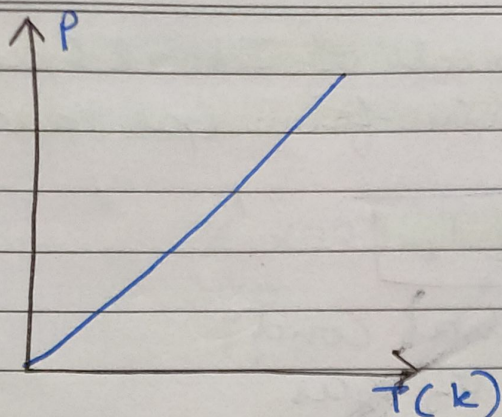
③ Amonton's or Gay-Lussac's law

$$V, n \rightarrow \text{const.}$$

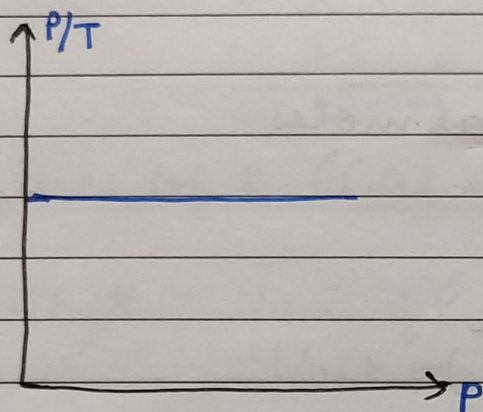
$$PV = nRT \quad (P \propto T) \quad T, P, n$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{P}{T} = \frac{nR}{V} \rightarrow \text{const. (K)}$$



Container \rightarrow Rigid
(egⁿ Cylinder)



(4) Avogadro's law \rightarrow "at same T , P equal volume of all gases contain equal no. of moles or molecules."

$P, T \rightarrow \text{const.}$

$V \propto n$

at STP / NTP \rightarrow 1 mole of any ideal gas occupy 22.4 L

1 bar \rightarrow 1 atm

at SATP (Std. Ambient $T \& P$) \rightarrow 1 mole of any ideal gas occupy 24.8 L

$\rightarrow 25^\circ\text{C} = T$
 $\rightarrow P = 1 \text{ bar}$

⊕ $PV = nRT \rightarrow$ Eqⁿ of state of Ideal Gas
(Obtain from Gas laws)

⊕ $P_1, V_1, T_1 \rightarrow P_2, V_2, T_2$
Initial condⁿ of Gas final condⁿ of gas

$n = \frac{PV}{RT_1}$

Initial = Final moles
moles

$$\frac{P_1 V_1}{RT_1} = \frac{P_2 V_2}{RT_2}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Q → How much volume of gas will occupy at STP when Gas at 300K, ~~2~~ 2 atm Pressure is brought to STP condⁿ?

Solⁿ ∴ $\frac{P_1 V_1}{T_1} = \left(\frac{P_2 V_2}{T_2} \right)_{\text{STP}}$

at STP

$T = 0^\circ\text{C}, 273\text{K}$

$P = 1\text{ bar} = 0.987\text{ atm}$

$2\text{ atm}, 300\text{K}, 2\text{L} \rightarrow 1\text{ bar}, 273\text{K}, V_2$

$$\frac{2 \times 2}{300} = \frac{1 \times V_2}{273}$$

$$\frac{4 \times 273}{300} = V_2$$

$$75$$

$$\boxed{\frac{273}{75} \text{ L} = V_2}$$

⊗ Values of R (Gas const.)

$$\text{Unit of } R = \frac{\text{Unit of } P \times \text{Unit of } V}{\text{Unit of } n \times \text{Unit of } T}$$

$$\left. \begin{aligned} R &= 0.0821 \text{ L-atm mol}^{-1} \text{ K}^{-1} \\ R &= 0.083 \text{ L-bar mol}^{-1} \text{ K}^{-1} \end{aligned} \right\} \rightarrow \approx \frac{1}{12}$$

$$R = 8.314 \text{ Joule mol}^{-1} \text{ K}^{-1}$$

$$R = 1.98 \text{ cal mol}^{-1} \text{ K}^{-1} \approx 2 \text{ cal mol}^{-1} \text{ K}^{-1}$$

$$\text{Data } P = 1 \text{ bar}$$

$$= 10^5 \text{ Pa} = 10^2 \text{ kPa}$$

$$V = 1 \text{ Litre} = 10^{-3} \text{ m}^3$$

$$\boxed{1 \text{ bar} \times 1 \text{ Litre} = 10^2 \times 10^{-3} (\text{kPa})(\text{m}^3)}$$

$$= 0.1 (\text{kPa})(\text{m}^3)$$

$$R = 0.083 \times (0.1 \text{ kPa})(\text{m}^3) \text{ mol}^{-1} \text{ K}^{-1}$$

$$= 0.0083 (\text{kPa})(\text{m}^3) \text{ mol}^{-1} \text{ K}^{-1}$$

$$= 8.314 \times 10^{-3} \text{ kPa m}^3 \text{ mol}^{-1} \text{ K}^{-1}$$

$$\textcircled{11} \quad PV = nRT$$

$$PV = \frac{m}{M} RT$$

$$PM = \frac{m}{V} RT$$

$$PM = dRT$$

M = molar mass of Gas

d = density of Gas

$$Mg = 2(V \cdot D)g$$

↓
vapour
density

$\textcircled{12}$ Boyle's law

$n, T \rightarrow \text{const.}$

$$PV = \text{const.}$$

$$P_1 V_1 = P_2 V_2 = nRT$$

$$\cancel{PM} = \cancel{d} \cancel{RT}$$

const. const.

$$P \propto d$$

$$P \uparrow, d \uparrow$$

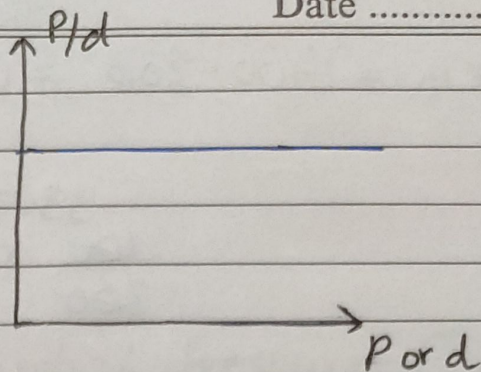
$$P = \left(\frac{RT}{M} \right) d$$

const. k

$$P = \text{const.}$$

Spiral d

$$\frac{P_1}{d_1} = \frac{P_2}{d_2}$$



(#) Charles's law

$$n, P \text{ const.}$$

$$PV = nRT$$

$$V = \frac{nRT}{P} \text{ const. (k)}$$

$$\frac{V}{T} = k$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{P}{\text{const.}} = \frac{Q}{R} \frac{1}{T} \text{ const.}$$

$$d \times T = \text{const.}$$

$$d_1 T_1 = d_2 T_2$$

Q → Gas at 300K with density of 2g/cc is heated such that density becomes 3g/cc. find % change in Temp. keeping Pressure const.?

Solⁿ

$$300 \times 2 = \frac{3}{2} T_2$$

$$\frac{300 \times 2 \times 2}{3} = 400 = T_2$$

$$\text{change in Temp.} = 400 - 300 = 100 \text{ K}$$

$$\frac{100}{300} \times 100 = 33.3\%$$

Q → An open vessel containing air at 300K is heated such that $\frac{3}{5}$ of the air is escaped. find Temp. at which heating was done?

Soln

$$n_1 \times T_1 = n_2 \times T_2$$

$$n_1 T_1 = \left(n_1 - \frac{3}{5} n_1 \right) T_2$$

$$n_1 T_1 = \frac{2}{5} n_1 T_2$$

$$\frac{150}{300 \times 5} = T_2$$

$$750 \text{ K} = T_2$$

⊕ open vessel Related Questions

for open vessel → (V, P) const.

$$\underbrace{(PV)}_{\text{const}} = n \underbrace{R}_{\text{const}} T$$

$$n \times T = \text{const.}$$

Date

$$\boxed{n_1/T_1 = n_2/T_2}$$

→ moles in container at Temp. T_2 .

→ moles in container at T_1

⊗ (#) Dalton's law of Partial Pressure (DLPP)

At const. T, V

→ "Total Pressure Exerted by mixture of non-reacting gases is Sum of Partial Pressure of each gas present in container."

egⁿ

Gas A	→ 1 atm
Gas B	→ 2 atm
Gas C	→ 3 atm

T, V const.

$$P_{\text{total}} = \underbrace{P_A}_{\text{Partial Pressure of A}} + \underbrace{P_B}_{\text{Partial pressure of B}} + \underbrace{P_C}_{\text{Partial Pressure of C}}$$

Partial Pressure of $\frac{z \text{ moles of gas}}{\text{Gas}}$ \times Total Pressure

$$= X_g \times P_T$$

egⁿ

300K

Gas A	→ 3 mole
Gas B	→ 4 mole

$P_{\text{Total}} = 5 \text{ atm}$

$P_A = ?$

$P_B = ?$
Spiral

$$P_A = \frac{3}{7} \times 5 = \boxed{\frac{15}{7}}$$

$$P_B = \frac{4}{7} \times 5 = \boxed{\frac{20}{7}}$$

$$\boxed{n_1 T_1 = n_2 T_2}$$

→ moles in container at T_1

→ moles in container at Temp. T_2

Dalton's law of Partial Pressure (DLPP)

At const. T, V

→ Total Pressure Exerted by mixture of non-reacting Gases is Sum of Partial Pressure of each Gas present in container."

egⁿ

Gas A → 1 atm
Gas B → 2 atm
Gas C → 3 atm

 $\Rightarrow P_{\text{total}} = P_A + P_B + P_C$

$T, V \text{ const.}$

Partial pressure of B

Partial Pressure of C

Partial Pressure of A

Partial Pressure of n moles of Gas \times Total Pressure

Gas Total moles

$$= X_g \times P_T$$

egⁿ 300K

Gas A → 3 mole
Gas B → 4 mole

$P_{\text{Total}} = 5 \text{ atm}$

$P_A = ?$

$P_B = ?$

Spiral

$$P_A = \frac{3}{7} \times 5 = \boxed{\frac{15}{7}}$$

$$P_B = \frac{4}{7} \times 5 = \boxed{\frac{20}{7}}$$

Q → $\boxed{\begin{matrix} O_2 = 32g \\ H_2 = 4g \end{matrix}}$ $P_{Total} = 10 \text{ atm}$
 $P_{O_2} = ?$
 $P_{H_2} = ?$

$$O_2 = 32 = \frac{32}{32} = 1 \text{ mole}$$

$$H_2 = \frac{4}{2} = 2 \text{ mole}$$

$$P_{O_2} = \frac{1}{3} \times 10 = \boxed{\frac{10}{3}}, P_{H_2} = \frac{2}{3} \times 10 = \boxed{\frac{20}{3}}$$

Q → Cylinder which can hold 2.82 L of water contains Gas at 20 atm, 300K. find no. of Ballons that can be filled at NTP if Diameter of each Ballon is 21 cm.

Solⁿ $V_1 = 2.82 \text{ L}$ $V_2 = ?$
 $T_1 = 300 \text{ K}$ \longrightarrow $T_2 = 273.15 \text{ K}$
 $P_1 = 20 \text{ atm}$ $P_2 = 1 \text{ atm}$

$$\frac{P_1 V_1}{T_1} = \frac{V_2 P_2}{T_2}$$

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}$$

~~$V_2 = \frac{20 \times 2.82 \times 273.15}{1 \times 300}$~~

~~$= 2.82 \times 273.15$~~

$$\frac{P_1 V_1 T_2}{T_1 P_2} = V_2$$

$$\frac{20 \times 2.82 \times 273}{300 \times 1} = 51.324 \text{ L} \Rightarrow 51324 \text{ ml}$$

(#) Vol. of Gas that can be filled at NTP

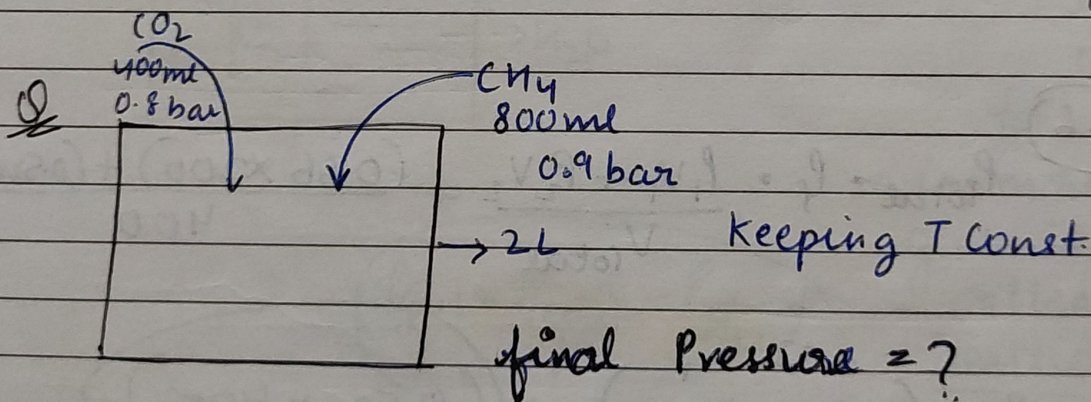
$$5132 - (2820) \text{ vol. of cylinder} \\ = 48504 \text{ ml}$$

$$V(\text{Ballon}) = \frac{4\pi r^3}{3} = \frac{4 \times \pi \times 21^3}{3} = \frac{4 \times \pi \times 21 \times 21 \times 21}{3}$$

$$= 11 \times 21 \times 21 = 4851 \text{ cm}^3$$

$$= 4851 \text{ ml}$$

(#) No. of Ballon that can be filled = $\frac{48504}{4851} = 10$
Ans



Solⁿ $P_f V_f = n_{\text{Total}} RT$

$$P_f \times 2 = (n_{\text{CO}_2} + n_{\text{CH}_4}) RT \Rightarrow P_f \times 2 = \left[\frac{P_{\text{CO}_2} V_{\text{CO}_2}}{RT} + \frac{P_{\text{CH}_4} V_{\text{CH}_4}}{RT} \right]$$

$$P_f = \frac{P_{\text{CO}_2} V_{\text{CO}_2} + P_{\text{CH}_4} V_{\text{CH}_4}}{2}$$

or

moles before mixing = moles after mixing

$$n_{\text{CO}_2} + n_{\text{CH}_4} = \underbrace{(n_{\text{CO}_2} + n_{\text{CH}_4})}_{\rightarrow n_{\text{total}}}$$

$$\frac{P_{\text{CO}_2} V_{\text{CO}_2}}{RT} + \frac{P_{\text{CH}_4} V_{\text{CH}_4}}{RT} = \frac{P_T V_{\text{container}}}{RT}$$

$$P_T = \frac{P_{\text{CO}_2} V_{\text{CO}_2} + P_{\text{CH}_4} V_{\text{CH}_4}}{V_{\text{container}}}$$

Step-2

(Q-6)

$$P_{\text{Total}} = P_f = \frac{P_1 V_1 + P_2 V_2}{V_{\text{Total}}} = \frac{(0.66 \times 200) + (0.52 \times 400)}{400}$$

$$P_T = \left(\overset{0.33}{\underbrace{0.66 \times \frac{200}{400}}} \right) + \left(\underbrace{0.52 \times \frac{400}{400}} \right)$$

(P_{CH_4}) in container
 \rightarrow Partial Pressure

$\rightarrow P_{\text{O}_2}$ in container

OR

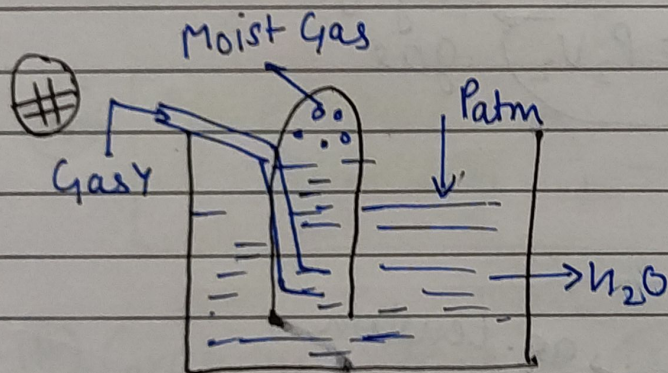
Apply Boyle's law for each Gasfor He

0.66 atm 200ml	$\xrightarrow[\text{const.}]{\text{mole \& T}}$	$P_{\text{He}} = P_2$ 400ml
-------------------	---	--------------------------------

$$P_1 V_1 = P_2 V_2$$

$$0.66 \times 200 = P_{\text{He}} \times 400$$

$$P_{\text{He}} = \frac{0.66 \times 200}{400} = \boxed{0.33}$$



$$P_{\text{moist gas}} = P_{\text{dry gas}} + P_{\text{H}_2\text{O vapour}}$$

Pressure due to water vapour in moist gas at given Temp. is called aqueous Tension

only depends on Temperature

$$P_{\text{gas}} = P_{\text{moist gas}} - P_{\text{H}_2\text{O}}$$

Q A closed container containing O_2 and some liquid water was found to exert 740 mm Pressure at $27^\circ C$.

i) Calculate the pressure exerted by O_2 if aqueous tension at $27^\circ C$ is 20 mm Hg.

Solⁿ $P_{O_2} = (740 - 20) = 720 \text{ mmHg} = \frac{720}{760} \text{ atm}$

ii) What will be final pressure when volume is doubled?

Solⁿ

$T = \text{const.}$

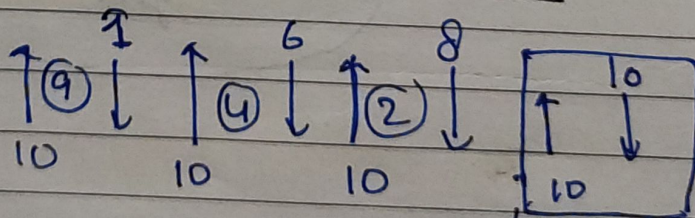
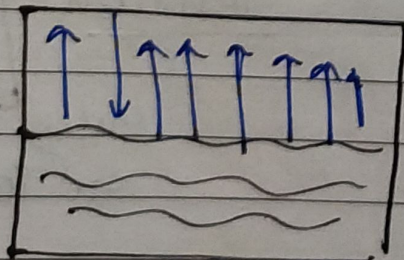
$P_1 V_1 = P_2 V_2$ only for gas

$720 \times V_1 = P_2 \times 2V_1$

$360 = P_2$

final Pressure $= 360 + \text{aq. tension}$
 $= 360 + 20 = 380 \text{ mmHg}$

Vapour Pressure of volatile liquid



Dynamic Eqⁿ

Rate of evaporation
=

Rate of condensatⁿ

at Given Temp
↓

Rate of evaporatⁿ
is fixed for given
surface of liq.

& space above liq. is
saturated with vapour

& pressure Exerted
by Saturated vapours
called Vapour Pressure.

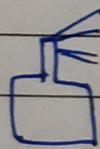
Depends on Temperature
only

(#) Vapour Pressure of liq. is independent of
(Volume)

- 1) Amount of liquid
- 2) Surface area of liquid
- 3) Size & shape of container.

(#)
$$\% \text{ Relative Humidity} = \frac{P_{H_2O} \text{ in moist air}}{V.P. \text{ of } H_2O} \times 100$$

(#) Graham's law of Diffusion and Effusion



→ Effusion

Ejection of
Gas at high
Pressure from
small orifice or
Pin hole.

↳ Spontaneous movement
of Gas particles / molecules
from Region of High Pressure
to low pressure.

Diameter of hole \geq Diameter of
molecule

" At a given T & P,
rate of Diffusion or Effusion is
inversely proportional to square root of
Vapour density of Gas or Molar mass of
Gas

$$\begin{aligned} r_{d/e} &\propto \frac{1}{\sqrt{V.D}} \\ r_{d/e} &\propto \frac{1}{\sqrt{M_g}} \end{aligned} \rightarrow \text{at given T \& P}$$

$$\textcircled{\#} r_{\text{effusion}} = \frac{P_g N_A A_{\text{hole}}}{\sqrt{2\pi M R T}} \rightarrow \text{Area of crosssection of hole}$$

$\textcircled{\#}$ for two gases Diffusing at Same T & P

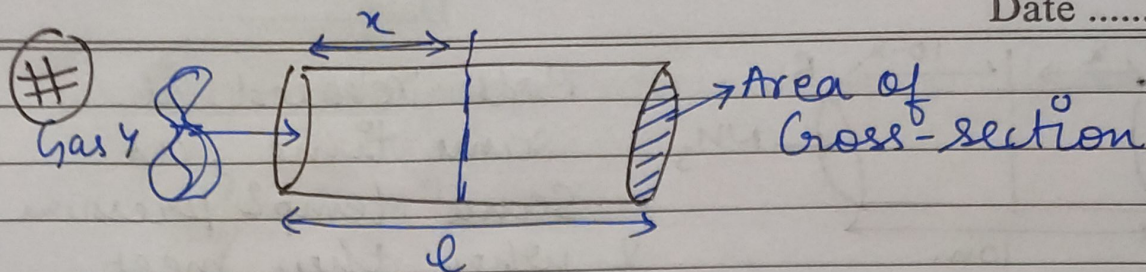
$$\frac{(r_d)_1}{(r_d)_2} = \sqrt{\frac{M_2}{M_1}}$$

$\textcircled{\#}$ At given T, But different pressures

$$r_{d/e} \propto \frac{P_g}{\sqrt{M}}$$

$$\frac{(r_{d/e})_1}{(r_{d/e})_2} = \frac{P_1}{P_2} \sqrt{\frac{M_2}{M_1}}$$

→ Rate of Diffusion or Effusion of Gas = $\frac{\text{moles or volume of Gas Diffused}}{\text{Time taken by Gas}}$



Volume of Gas Diffused upto distance x
 $= x \times A$

Volume of gas diffused = Area of cross-section of tube \times Distance travelled by Gas

\Rightarrow 30 ml (O_2) diffused in 5 mins at 3 atm
 find Vol. of (SO_2) diffused in same time
 at 2 atm at same temp.?

$$r_{O_2} \propto \frac{P_{O_2}}{\sqrt{M_{O_2}}}$$

$$\frac{r_{O_2}}{r_{SO_2}} = \frac{P_{O_2}}{P_{SO_2}} \sqrt{\frac{M_{SO_2}}{M_{O_2}}}$$

$$r_{SO_2} \propto \frac{P_{SO_2}}{\sqrt{M_{SO_2}}}$$

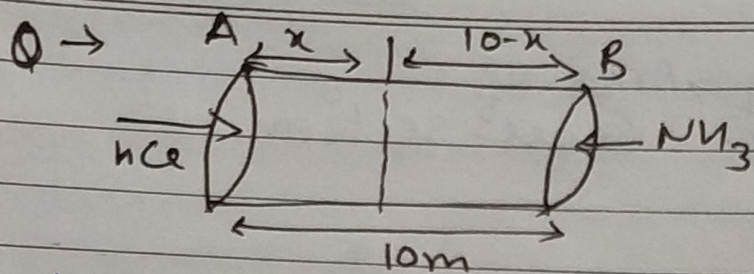
$$\frac{V_{O_2}/t}{V_{SO_2}/t} = \frac{3}{2} \sqrt{\frac{64}{32}}$$

$$\frac{V_{O_2}}{V_{SO_2}} = \frac{3 \times 8}{2 \times \sqrt{2}} = \frac{3}{\sqrt{2}}$$

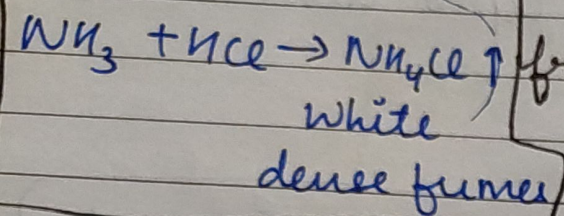
$$\frac{30 \times \sqrt{2}}{3} = V_{SO_2}$$

Spiral

$$V_{SO_2} = 10\sqrt{2} \text{ ml}$$



Note :-



Both released at same time, at same temp & pressure & when they meet white dense fumes are observed. find at what distance from terminal A, fumes are observed?

Solⁿ

$$\frac{r_{\text{NH}_3}}{r_{\text{HCl}}} = \sqrt{\frac{M_{\text{HCl}}}{M_{\text{NH}_3}}} = \sqrt{\frac{36.5}{17}}$$

$$\frac{v_{\text{NH}_3} / t}{v_{\text{HCl}} / t} \approx \sqrt{2} \approx 1.5$$

$$\frac{(10-x) \times A}{x \times A} = \frac{3}{2}$$

$$20 - 2x = 3x$$

$$20 = 5x$$

$$4 = x$$

(Q) Sulphur vapours travels 2 times slower than oxygen gas. (S_n)
find n ?

Solⁿ
$$\frac{(rd)_{S_n}}{(rd)_{O_2}} = \frac{1}{2} = \sqrt{\frac{M_{O_2}}{M_{S_n}}} = \sqrt{\frac{32}{32n}}$$

$$\frac{1}{2} = \sqrt{\frac{1}{n}}$$

$$\frac{1}{4} = \frac{1}{n}$$

$$\boxed{n = 4}$$

* Kinetic theory of Gases (for Ideal Gases only)

- 1) Gas molecules are considered as point masses.
- 2) Vol. of ~~volume~~ 1 molecule is negligible compared to Total vol. of container.

$\boxed{\text{Vol. of Gas} = \text{Vol. of container}}$

- 3) No attraction and no repulsion b/w gas molecules.
- 4) Collision b/w particles (molecules) will be perfectly elastic. There will be no loss of energy.
Momentum conserved.

- 5) Average kinetic energy or average speed of Gas molecules is only temp. dependent. (kelvin)

$$\overline{K.E} \text{ or } (K.E)_{av} \text{ or } (K.E)_{Total} \propto T$$

- 6) Effect on molecular motion due to gravity is ignored or negligible.

7) Kinetic Gas eqn $\rightarrow PV = \frac{1}{3} m N u_{rms}^2$

(P) is Pressure of Gas

(V) is Volume of Gas or Container

(m) is mass of 1 molecule of Gas

(N) is no. of Total molecules

(u_{rms}) is root mean square speed

$$u_{rms} = \sqrt{\overline{u^2}}$$

$u_{rms}^2 = \text{mean square speed}$
($\overline{u^2}$)

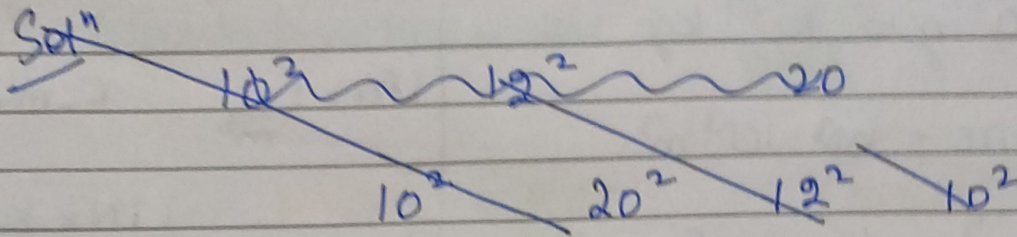
Q \rightarrow

Gas Y

40 molecules	$\rightarrow 10 \text{ m/s}$
20 molecules	$\rightarrow 12 \text{ m/s}$
10 molecules	$\rightarrow 20 \text{ m/s}$
30 molecules	$\rightarrow 13 \text{ m/s}$

① u_{rms} of Gas Y?

② u_{av} (average speed)? ③ u_{mp} (most probable speed)?



~~u_{rms}~~ =

$$(i) \quad u_{rms} = \sqrt{\frac{30 \times (13)^2 + 10 \times (20)^2 + 20 \times (12)^2 + 40 \times (10)^2}{100}} = \sqrt{\frac{N_1 u_1^2 + N_2 u_2^2 + \dots}{N_1 + N_2 + \dots}}$$

(ii) u_{mp} = speed attained by max. fraction of molecules in given sample at any temp.

$$= 10 \text{ m/s}$$

$$(ii) \quad u_{av} = \frac{(10 \times 40) + (12 \times 20) + (20 \times 10) + (30 \times 13)}{100}$$

$$= \frac{N_1 u_1 + N_2 u_2 + N_3 u_3 + \dots}{N_1 + N_2 + N_3 + \dots}$$

Q → Calculate Pressure in Pascals Exerted by 10^{23} molecules each having 10^{-22} grams in a container of Volume 1 litre. The rms speed is 10^5 cm/sec. Also find Total K.E of these particles? find Temp. of Gas?

Solⁿ $m = 10^{-22} \text{ g} = 10^{-25} \text{ kg}$

$V = 1 \text{ Litre} = 10^{-3} \text{ m}^3$

$u_{\text{rms}} = 10^5 \text{ cm/s} = 10^5 \times 10^{-2} \text{ m/s}$
 $\approx 10^3 \text{ m/s}$

$PV = \frac{1}{3} m N u_{\text{rms}}^2$

$1 \text{ Pa} = \frac{1 \text{ N}}{\text{m}^2}$ $F = m \cdot a$

$1 \text{ N} = \frac{\text{kgm}}{\text{s}^2}$

$1 \text{ m}^3 = 10^3 \text{ L}$

$10^{-3} \text{ m}^3 = 1 \text{ L}$

$1 \text{ Pa} = \frac{\text{kg}}{\text{m s}^2}$

$P \times 10^{-3} = \frac{1}{3} \times 10^{-25} \times 10^{23} \times 10^6$

$P = \frac{10^7}{3} \text{ Pa}$

$P = \frac{10^7}{3 \times 1.01325 \times 10^5} \text{ atm}$

$3 \times 1.01325 \times 10^5$

$1 \text{ bar} = 10^5 \text{ Pa} = 10^5 \frac{\text{N}}{\text{m}^2}$

$1 \text{ atm} = 1.01325 \times 10^5 \text{ Pa}$

$\overline{K \cdot E} = ?$

$PV = \frac{1}{3} N m u_{\text{rms}}^2$

$PV = \left(\frac{N}{N_A} \right) RT$ $\rightarrow \text{moles}$

$\frac{3}{2} PV = \left(\frac{1}{2} m N u_{\text{rms}}^2 \right)$

$\rightarrow (K \cdot E)_{\text{Total}}$

$\frac{3}{2} \frac{N}{N_A} RT = (K \cdot E)_{\text{Total}}$

⊕ $(\overline{K.E})$ for 1 molecule $\rightarrow \frac{3RT}{2N_A}$ $(N=1)$

$(\overline{K.E})$ for 1 mole $= \frac{3RT}{2}$ $(N=N_A) \rightarrow \text{mole}$

$(\overline{K.E})$ for n moles of any Ideal Gas $= \frac{3}{2} nRT$

$\overline{K.E}/\text{molecule} = \frac{3RT}{2N_A} = 3KT$

$K = \frac{R}{N_A} = \frac{8.314 \text{ J mol}^{-1} \text{ K}^{-1}}{6.023 \times 10^{23} \text{ particle mol}^{-1}} = 1.38 \times 10^{-23} \text{ J K}^{-1}$
 \rightarrow Boltzmann constant

Q $\overline{K.E}$ for 5 He gas atom at 300K?

$\overline{K.E} = \left(\frac{3}{2} NKT = \frac{3}{2} \times 5 \times 1.38 \times 10^{-23} \times 300 \right) \text{ Joules}$

⊕ Different type of speeds for sample of Gas molecules

① Root mean Square speed (u_{rms})

$PV = \frac{1}{3} mN u_{rms}^2$

for 1 mole ; $N = N_A$

$$m N_A = \text{Molar Mass (M)}$$

$$PV = \frac{1}{3} M u_{rms}^2$$

$$PM = dRT$$

$$PV = RT \rightarrow \text{for 1 mole}$$

$$3PV = M u_{rms}^2$$

$$u_{rms} = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3P}{d}}$$

only Temp. dependent

$$(2) \quad u_{av} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8PV}{\pi M}} = \sqrt{\frac{8P}{\pi d}}$$

$$\frac{8}{\pi} \approx 2.54$$

$$\pi = 3.14 \text{ or } 22/7$$

$$(3) \quad u_{mp} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2PV}{M}} = \sqrt{\frac{2P}{d}}$$

Q → find u_{mp} for O_2 gas at 300K?
in (m/s)

$$= \sqrt{\frac{2 \times 8.314 \times 300}{32 \times 10^{-3}}}$$

$$R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$$

$$= 80.0821 \frac{\text{L-atm}}{\text{mol K}^{-1}}$$

$$1 \text{ L-atm} = 101.325 \text{ Joules}$$

(11) $u_{mp} < u_{av} < u_{rms}$

(12) $u_{mp} : u_{av} : u_{rms}$

$$\sqrt{2} : \sqrt{8/\pi} : \sqrt{3}$$

$$1 : 1.12 : 1.22$$

$$0.816 : 0.921 : 1$$

Q for a gas

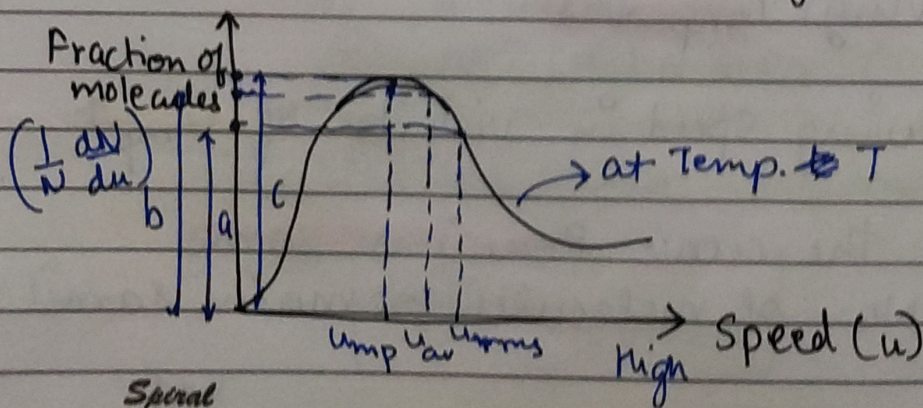
$$u_{mp} = x \text{ m/s}$$

$$u_{rms} = ? \quad u_{av} = ?$$

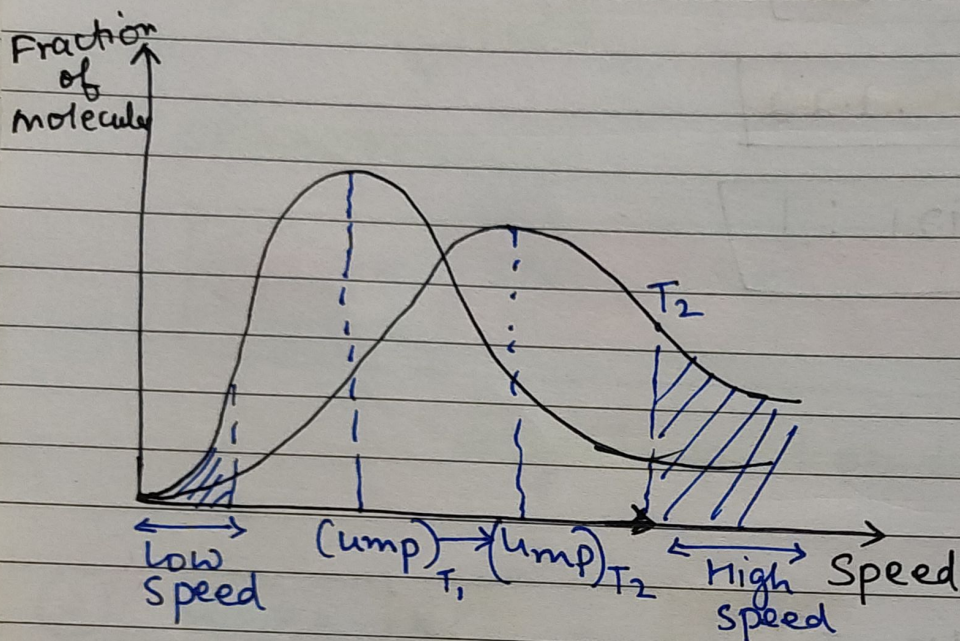
Solⁿ $u_{rms} = \frac{1.22 u_{mp}}{1.22 x}$

$$u_{av} = 1.12 x$$

(#) Boltzmann - Distribution curve for distribution of molecular speeds at a given temp.



$a < b < c$
 fraction having u_{rms} fraction of molecules having speed = u_{mp}
 fraction having u_{av}



$T_2 > T_1$ on \uparrow ing Temp.

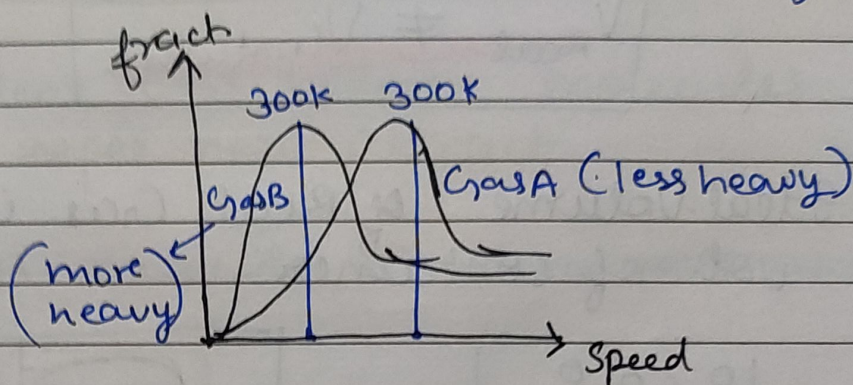
Pg-48, 51, 52

- 1) Fraction of molecules having low speed \downarrow es
- 2) Fractⁿ of molecules having high speed \uparrow es.
- 3) Value of u_{mp} , u_{rms} , u_{av} \uparrow
- 4) Fraction having u_{mp} \downarrow es
- 5) Molecules having speed in vicinity of u_{mp} \uparrow
- 6) Area under the curve Remains same
(as Total No. of molecules remains same)

7) At same temp.

Heavy gas have low ump value \rightarrow more fraction
 lighter gas " high ump value \rightarrow less fraction

$$u_{mp} \propto \frac{1}{\sqrt{m}}$$



• Real Gas (Existing gas, Practical gas)

$$\rightarrow P, V \neq nRT$$

$$P_i V_i = nRT$$

$$(P_R)_I (V_R)_I = nRT$$

\rightarrow Don't follow Gas laws

\rightarrow Two postulates of K.T.G (Kinetic Theory of Gas) not obeyed by Real Gas. Real Gas Behaviour deviates from Ideal Gas behaviour.

Postulate 1 \rightarrow Vol. of single molecule of ideal Gas is negligible w.r.t total vol. of container at all T & P, so total vol. of container is available for free movement of Gas which is called as Ideal volume.

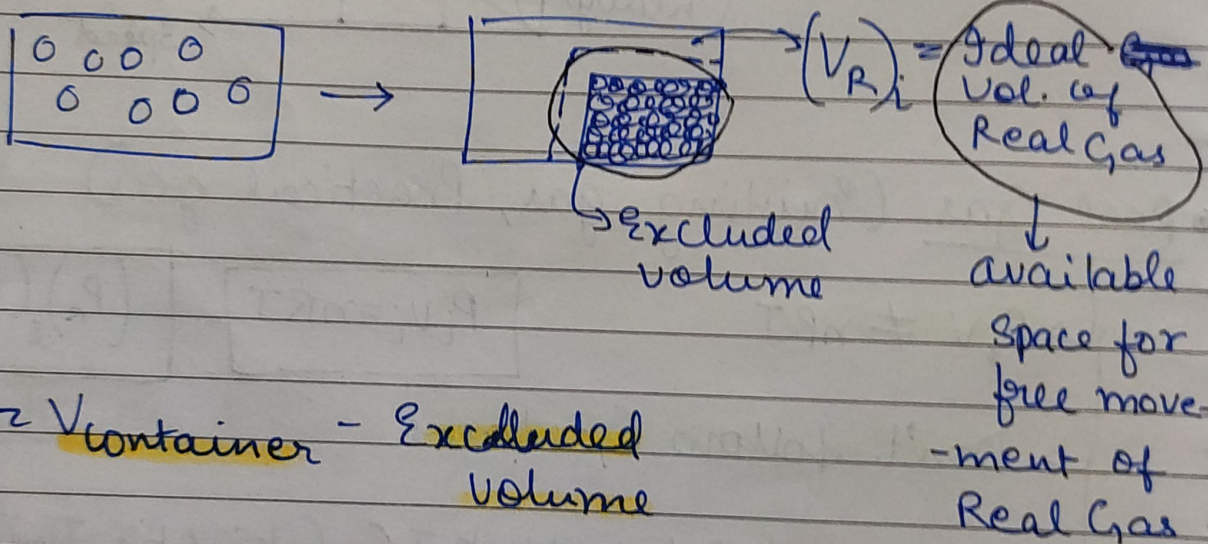
$$V_i = V_{\text{container}}$$

But for Real Gas, at high Pressure (Vol. of container is small) size for real Gas molecule is significant & cannot be ignored.

So,

$$V_{\text{Real}} \neq V_{\text{Container}}$$

Ideal Volume of Real Gas is different from vol. of container.



$$\approx V_{\text{Container}} - \text{Excluded Volume}$$

$$(V_R)_i = V_{\text{Container}} - \text{vol. correction term}$$

Excluded volume for single molecule = $4 \times \text{Vol. of molecule}$
 $\approx 4 \times \frac{4}{3} \pi r^3$

Excl. vol. for 1 mole Gas $\approx 4 N_A V_{\text{molecule}}$
 $\approx b$ (co-volume of Gas)

Excluded Vol. for n moles of Gas = nb

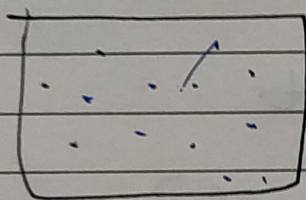
So,
$$(V_p)_I = V_{\text{container}} - nb = V - nb$$

Postulate 2: Ideal Gas ~~volume~~ molecules neither repel nor attract.

But Real Gas molecules repel as well as attract

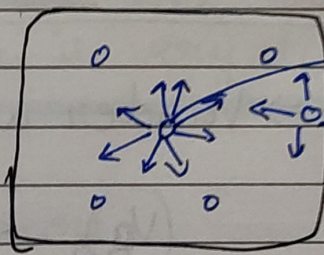
at Moderate Pressure

Force of attrⁿ b/w particles dominate



Ideal Gas

Press. = P_i



P_R

$F_{\text{net}} = 0$
 $F_{\text{net}} \neq 0$

(This molecule collides with less force compared to molecule of Ideal Gas are FOA)

$$(P_R)_I = P_R + P_{\text{correction}}$$

Ideal Pressure of Real Gas

$$P_{\text{correction}} \propto \left(\frac{n}{V}\right) \left(\frac{n}{V}\right)$$

loss of Pressure due to FOA.

moles of per unit vol. attracting
moles per unit volume of container striking the wall of container.

$$(P_R)_i = P_R + \frac{an^2}{V^2}$$

$$P_{\text{correction}} = \frac{an^2}{V^2}$$

$a \rightarrow$ vander waal constant
 \rightarrow Measure of force of attrⁿ
 \rightarrow Different for different gases

(#) Vander waal Eqⁿ

$$(P_R)_i (V_R)_i = nRT$$

$$\left(P + \frac{an^2}{V^2} \right) (V - nb) = nRT$$

$a, b \Rightarrow$ Vander waal const.

$$P_R < P_i$$

$$(V_R)_i < V_{\text{container}} (V)$$

$$(P_R)_i = P_R + \frac{n^2 a}{V^2}$$

$$(V_R)_i = V - nb$$

$$\left(P + \frac{n^2 a}{V^2} \right) (V - nb) = nRT$$

for 1 mole, $\left(P + \frac{a}{V^2} \right) (V - b) = RT$

(#) value of a & b different for different Gases

(#) for any Gas, value of $a \gg \gg b$.

(#) Value of $a \propto$ Size ~~of~~ & mass of Gas molecule.

Value of $b \propto$ Size & mass of Gas molecule.

(*) value of a & b
 $\boxed{CH_4 < SO_2}$

(#) $a \approx 0 \longrightarrow$ molecules are very small (He, H_2)
 and or

Pressure is very low (Size of container is very large, so that distance b/w molecule very large)

(#) $b \approx 0 \longrightarrow$ Moderate Pressure
 or
 low Pressure

(#) $a \neq 0 \longrightarrow$ at Mod. Pressure

(#) $a = 0 \longrightarrow$ at High Pressure (because it feels repulsion)

(#) $b \neq 0 \longrightarrow$ at high Pressure

(#) $P_{\text{correction}} = \frac{a n^2}{V^2} \longrightarrow \frac{\text{mol}^2}{L^2 \text{ or } (dm^3)^2}$
 Spiral \rightarrow atm

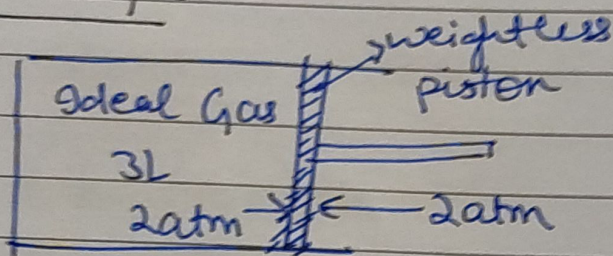
Unit of $a = \text{atm} \cdot \text{dm}^6 \cdot \text{mol}^{-2}$

Unit of $b = \text{L/mol}$



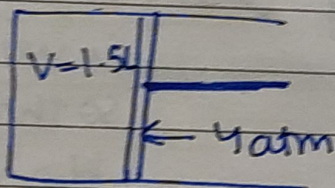
Compressibility factor (Z) = $\frac{V_R \text{ at } T, P}{V_i \text{ at } T, P}$

Ideal Gas



$$Z = \frac{V_R}{\frac{nRT}{P}} = \frac{PV_R}{nRT}$$

\downarrow
P ↑ by 2 times
 $T = \text{const.}$

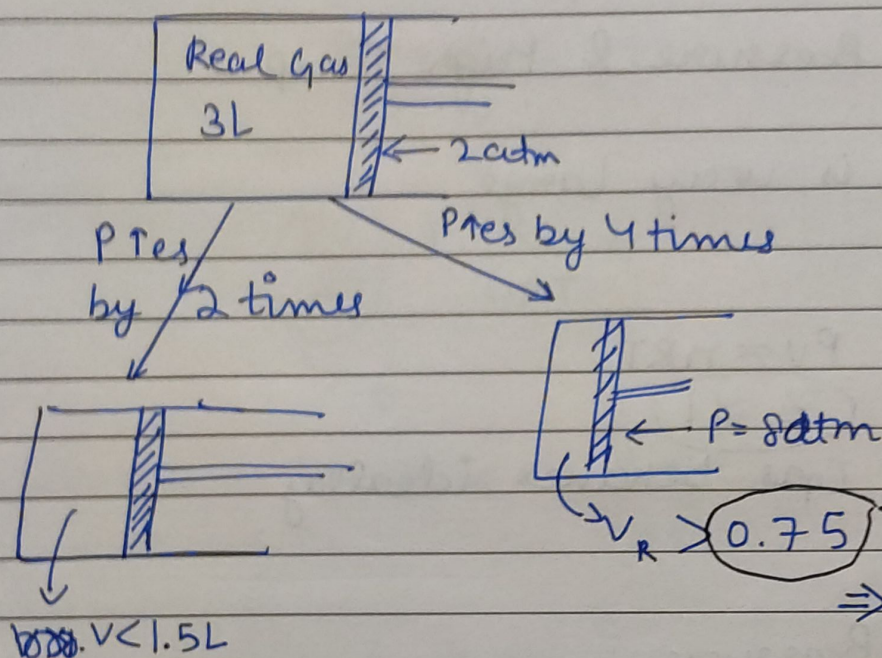


$$P_1 V_1 = P_2 V_2$$

$$V_2 = \frac{P_1 V_1}{P_2} = \frac{2 \times 3}{4} = \boxed{1.5 \text{ L}}$$

Compressibility of Ideal gas
is according to Boyle's law

$$\left(P \propto \frac{1}{V} \right) \text{ at given } T \text{ and fixed } n.$$

Real Gas

(A) less compressible
 $V_R = 0.83 L$
 $V_R = 0.79 L$
 (B)

$$Z > 1$$

\Rightarrow Compress. < Ideal Gas
 of Real Gas

$$P_1 V_1 \neq P_2 V_2$$

$$P V_R < nRT$$

$$\Rightarrow P V_i = nRT$$

$$\Rightarrow P V_R > nRT$$

co-volume
 (b)

\Rightarrow Force of repulsion dominates b/w molecules

\Rightarrow More is volume of Real Gas greater than Vol. of ideal Gas, ~~more~~ will be its

\Rightarrow compressibility of Real Gas > Compressibility of Ideal Gas

\Rightarrow Force of attractⁿ exist b/w molecules.

$$\Rightarrow P V_R < nRT, \frac{P V_R}{nRT} < 1$$

$$Z < 1$$

$$\textcircled{\#} Z = 1 \quad (\text{for ideal Gas})$$

Vander waal eqⁿ at different conditions

① At very low Pressure & High Temp.

Vol. of cont. is very large

$$\begin{matrix} a \approx 0 \\ b \approx 0 \end{matrix}$$

$$PV \approx nRT$$

$$Z \approx 1$$

Real Gas behaves ideally

② At moderate Pressure

for SO_2

$$a \approx 6.5$$

$$b = 0.03$$

$$V - \underbrace{(nb)}_{\approx 0} \approx V$$

##

Size of container

& Distance b/w molecules is still large enough. So, that $b \approx 0$

But $a \neq 0$ (Since at mod. Pressure a dominates)

$$\left(P + \frac{an^2}{V^2} \right) V = nRT$$

for 1 mole,

$$PV + \frac{a}{V} = RT$$

$$\left(\frac{PV}{RT} \right) + \frac{a}{VRT} = 1$$

\searrow
 z

$$z = 1 - \frac{a}{VRT}$$

$$\boxed{z < 1}$$

③ at high Pressure

$$a = 0$$

$$b \neq 0$$

for 1 mole,

$$P(V-b) = RT$$

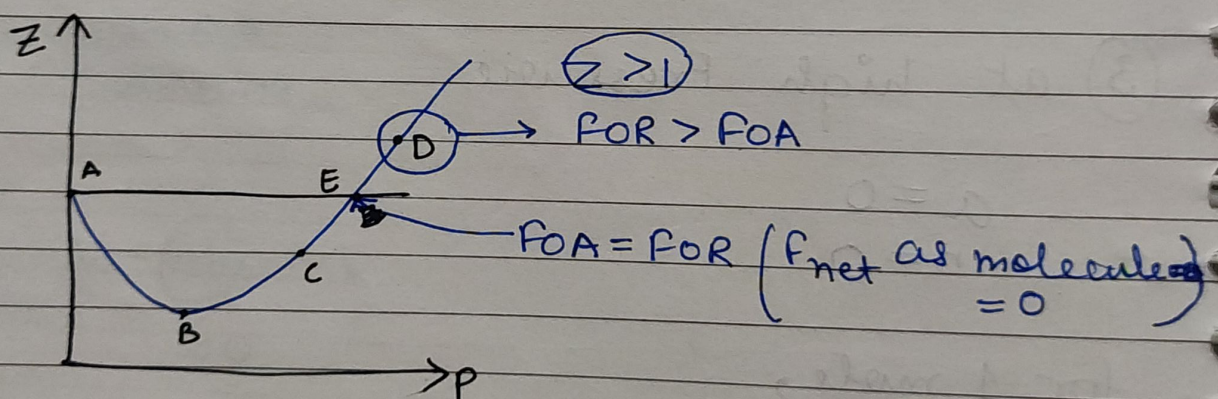
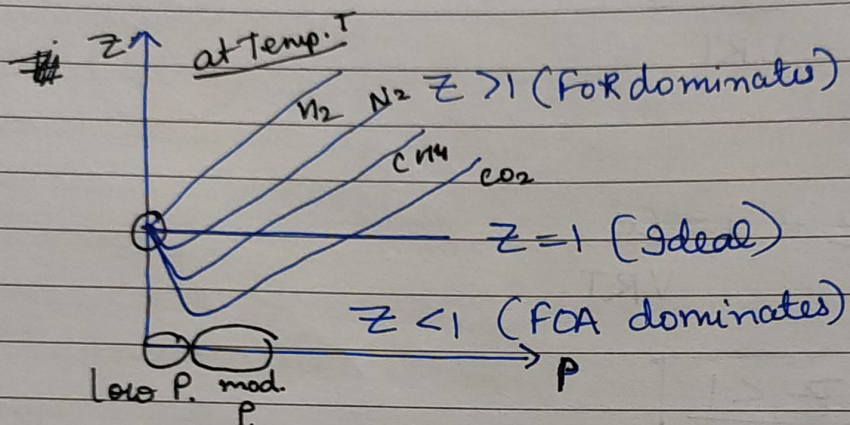
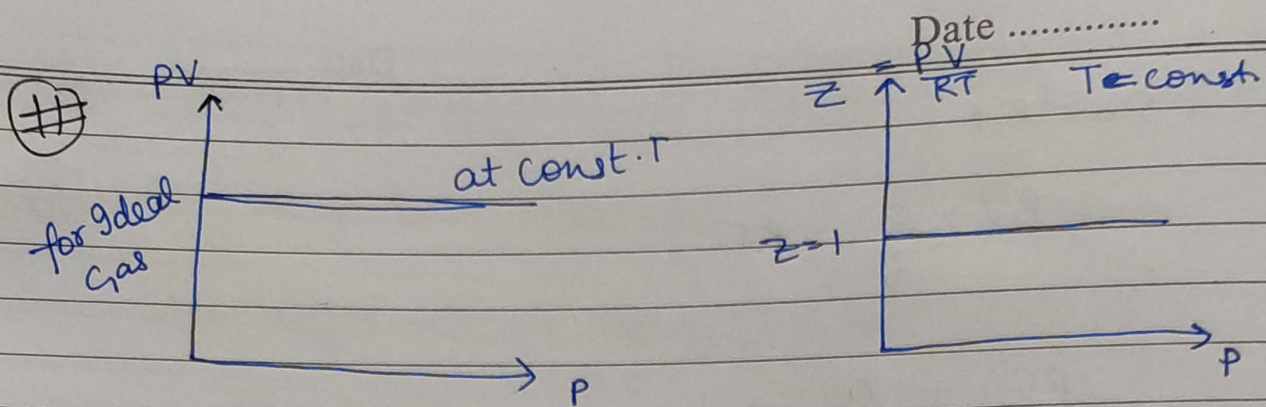
$$PV - Pb = RT$$

$$\frac{PV}{RT} - \frac{Pb}{RT} = 1$$

$$z = 1 + \frac{Pb}{RT}$$

$$\boxed{z > 1}$$

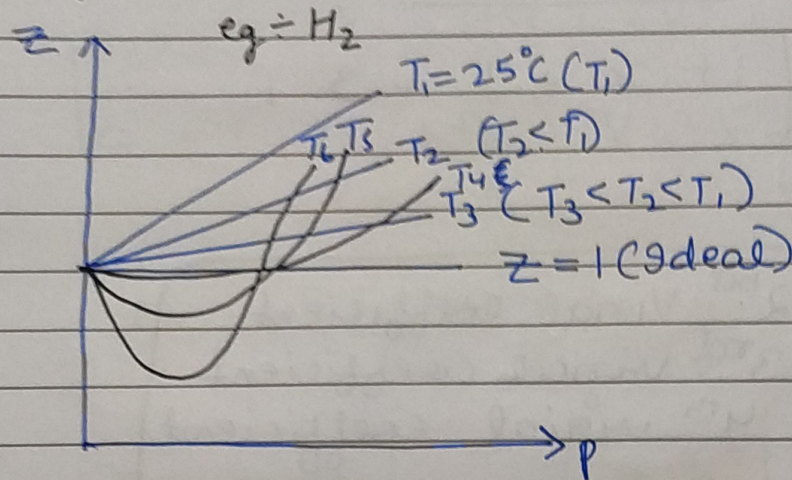
Spiral



- A to B, $Z \downarrow$, $FOA \gg FOR$
- at B, $Z \text{ min}$, $FOA \rightarrow \text{max}$
- at C, attrⁿ & Replⁿ both exist
But Total attrⁿ > Replⁿ

For He, H₂ → only FOR exist
at room temp.
 $Z > 1$

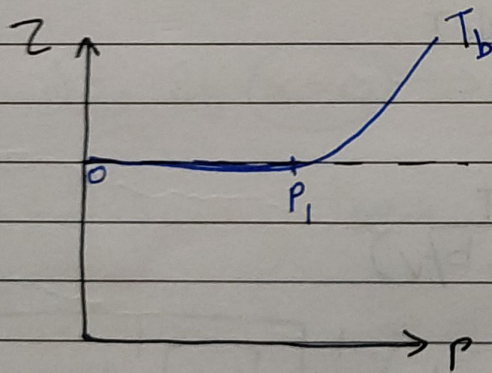
for any gas



$$T_6 < T_5 < T_4 < T_3 < T_2 < T_1$$

⑧ Boyle's Temp. of Gas (T_b)

Temp. at which real gas behaves ideally upto ~~some~~ ^{certain} pressure range. ~~certain~~



slope

$$\left(\frac{dz}{dp} \right)_{T_b} = 0 \quad \left(\text{from } P=0 \text{ to } P=p_i \right)$$

$$\left[\frac{dPV}{dP} \right]_{T_b} = 0 \quad \left(T_b > \frac{a}{b} \right) \quad T_b = \frac{a}{Rb}$$

• If T_{gas} is greater than its Boyle's Temp.

Z res, Compressibility of Real Gas \downarrow

$$V_R > V_I$$

(11) Virial equation for Real Gas

$$PV = RT \left[1 + \frac{x}{V} + \frac{y}{V^2} + \frac{z}{V^3} + \dots \right]$$

(for 1 mole)

$x \rightarrow 2^{\text{nd}}$ Virial coefficient
 $y \rightarrow 3^{\text{rd}}$ virial coefficient
 $z \rightarrow 4^{\text{th}}$ virial coefficient.

Derivatⁿ

$$\left(P + \frac{a}{V^2} \right) (V - b) = RT \quad (\text{for mole})$$

$$\left(P + \frac{a}{V^2} \right) V \left(1 - \frac{b}{V} \right) = RT$$

$$\left(PV + \frac{a}{V} \right) \left(1 - \frac{b}{V} \right) = RT$$

$$\left(PV + \frac{a}{V} \right) = \frac{RT}{\left(1 - b/V \right)}$$

$$\left(PV + \frac{a}{V} \right) = RT \left(1 - \frac{b}{V} \right)^{-1}$$

$$\left[1 - x \right]^{-1} = 1 + x + x^2 + x^3 + \dots$$

$x \ll 1$

$$PV + \frac{a}{V} = RT \left[1 + \frac{b}{V} + \frac{b^2}{V^2} + \dots \right]$$

$$PV = RT \left[1 + \frac{b}{V} - \frac{a}{VRT} + \frac{b^2}{V^2} + \dots \right]$$

Spiral

$$\left(1 - \frac{b}{V} \right)^{-1} = 1 + \frac{b}{V} + \frac{b^2}{V^2} + \frac{b^3}{V^3} + \dots$$

$$PV = RT \left[1 + \frac{1}{V} \left(\frac{b-a}{RT} \right) + \frac{b^2}{V^2} + \frac{b^3}{V^3} + \dots \right]$$

$$Z = 1 + \frac{1}{V} \left(\frac{b-a}{RT} \right) + \frac{b^2}{V^2} + \frac{b^3}{V^3} + \dots$$

$$PV = RT \left[1 + \frac{x}{V} + \frac{y}{V^2} + \frac{z}{V^3} + \dots \right]$$

$$x = b - \frac{a}{RT}$$

$$y = b^2$$

$$z = b^3$$

at $T_b \Rightarrow PV \approx RT$ Gas behaves ideally

$$\frac{1}{V^2}, \frac{1}{V^3} + \dots \text{negligible}$$

So, If $(x=0) \rightarrow$ Second virial coeff.

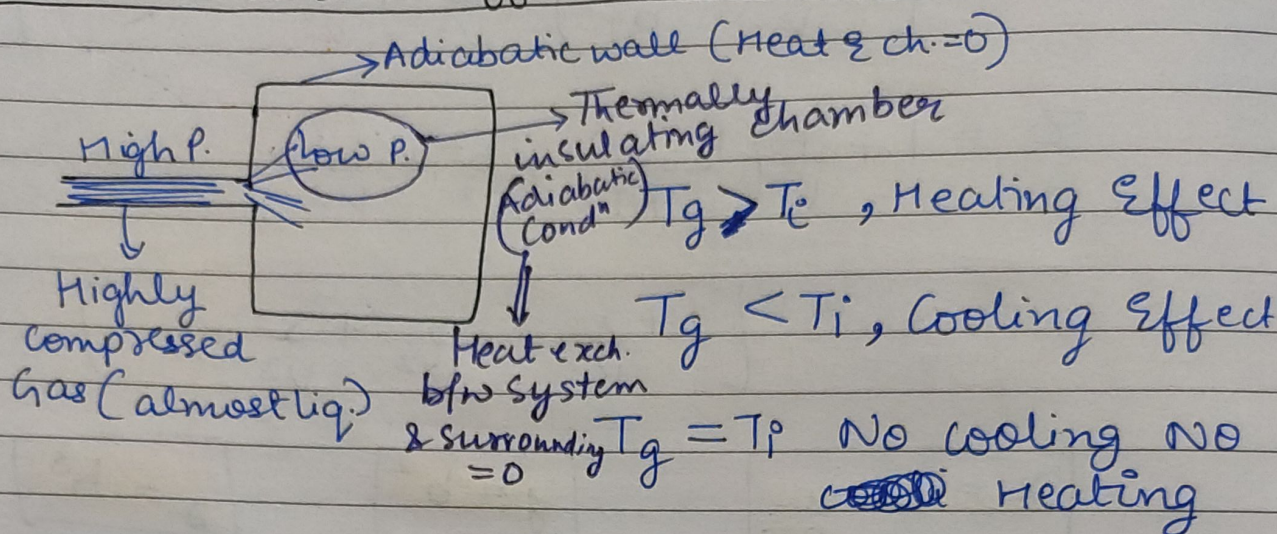
then $PV \approx RT$ & Temp. is called (T_b)

$$x = \frac{b-a}{RT_b} = 0$$

$$\frac{a}{RT_b} = b$$

$$T_b = \frac{a}{Rb}$$

Joule - Thomson Effect & Inversion effect



$$T_i = 2T_b$$

$$T_i = 2 \left(\frac{a}{Rb} \right)$$

for any gas,

$$T_i > T_b > T_c$$

If Temp. of gas during expansion from High P. to Low P. is below its inversion temp, then cooling effect is ^{observed} [Adiabatic cond'n]

Liquification of Real Gases

Real Gas behaves ideally at low P, high T $\Rightarrow a=0$
 $b=0$

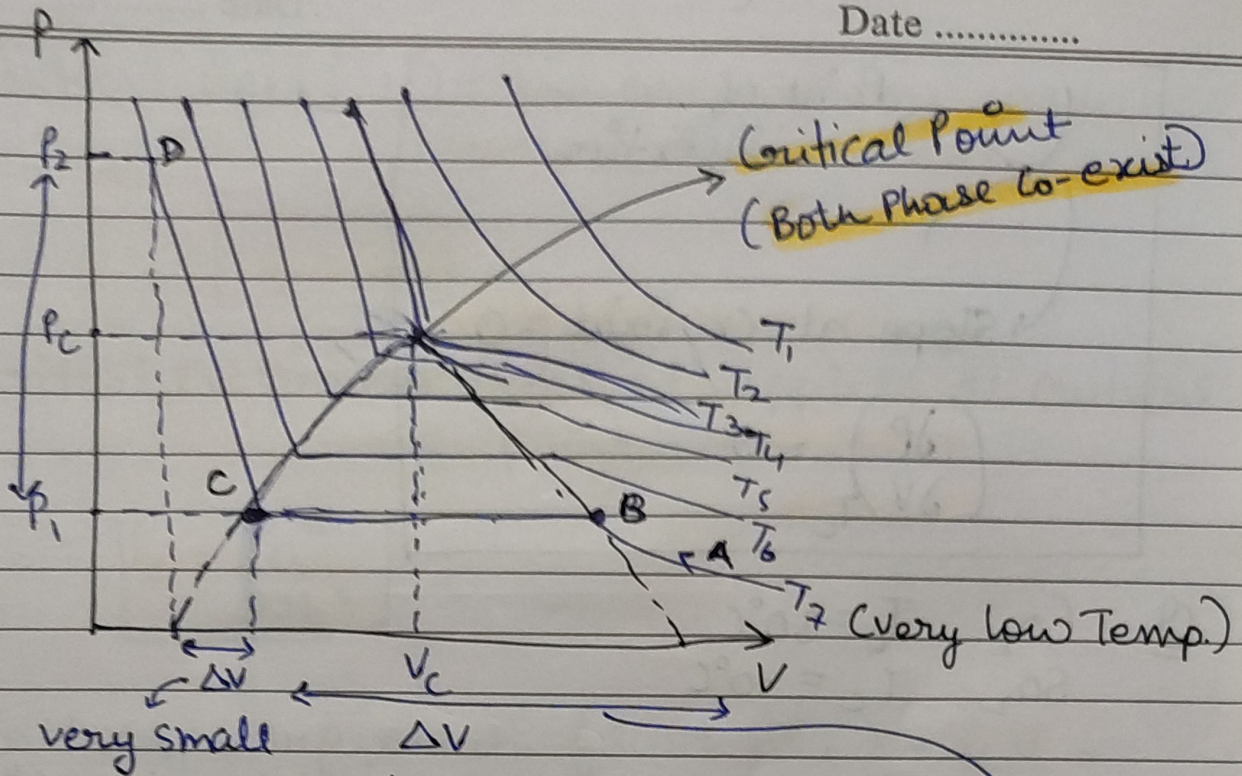
Ideal Gases cannot be liquified.

$$P \approx nRT$$

$$Z \approx 1$$

Req. conditions \Rightarrow High Pressure \rightarrow Bring Molecules closer to each other.

Low Temperature \rightarrow decreases K.E of gas so that liquification becomes easy.

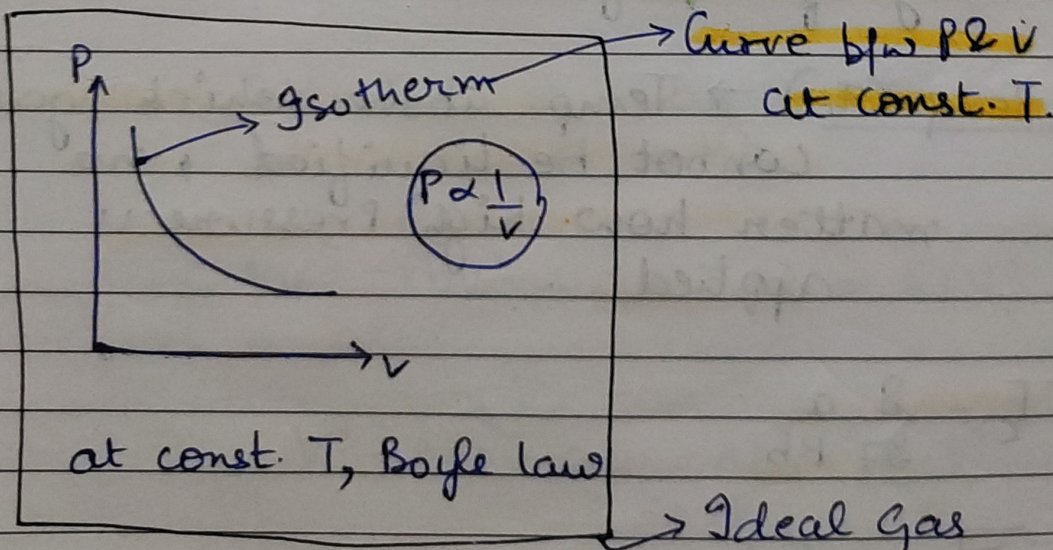


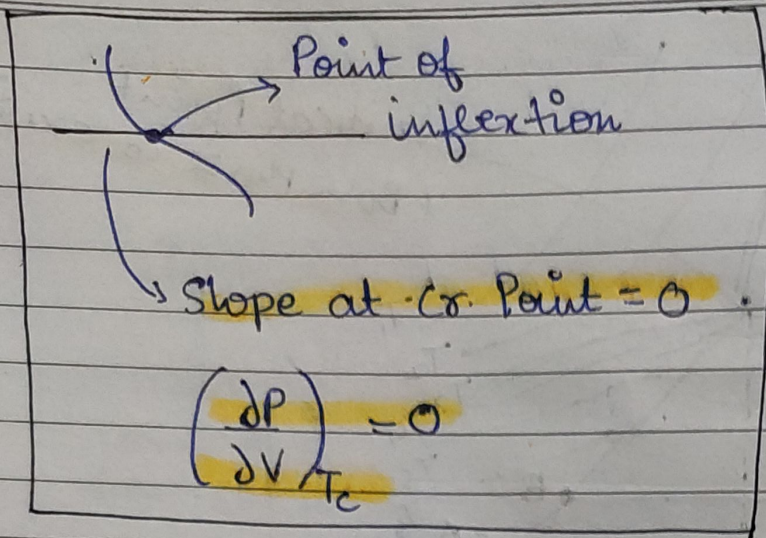
Andrews Isotherm Curve

- AB → gas phase
- at B → gas
- B to C → $g \rightleftharpoons g$
- at C → liq.
- CD → liq. phase

large Vol. change
↓
suggesting
gas ⇌ liq.

- B to C phase change occurs at const. P & T.





$$\begin{array}{ll} \text{Co}_2 & T_c = 30^\circ\text{C} \\ \text{So}_2 & T_c = 40^\circ\text{C} \end{array}$$

Which is having more tendency of liquification at 15°C ?

Solⁿ So_2

- # For liquification, Temp. is most important factor.
- # More is value of T_c for any gas, more will be its tendency of liquification.
- # Critical Temp. (T_c) \rightarrow Temp. above which gas cannot be liquified, no matter how high Pressure is applied.

$$T_c = \frac{8a}{27Rb}$$

Critical Temp. of Gas $\left[T_c = \frac{8}{27} (T_b) \right] \rightarrow$ Boyle Temp. of Gas

⊕ Critical Vol. (V_c) = Vol. occupied by gas at critical T & critical P.

$$V_c = 3b$$

⊕ Critical Pressure \Rightarrow Pressure applied at critical point.

$$P_c = \frac{a}{27b^2}$$

⊕ Critical compression factor (Z_c) = $\frac{P_c V_c}{RT_c} =$
 $= \frac{a}{9 \cdot 27b^2} \times 3b = \frac{a}{9b} / \frac{8a}{27Rb} \times R = \frac{3}{8} \frac{RT_c}{RT_c} =$

$\frac{3}{8}$ for vander waal Gas.

for 1 mole

$$\left(P + \frac{a}{V^2}\right)(V - b) = RT$$

$$PV - Pb - \frac{a}{V} - \frac{ab}{V^2} = RT$$

$$PV^3 - PbV^2 + aV - ab = RTV^2$$

$$V^3 - bV^2 - \frac{aV}{P} - \frac{ab}{P} = \frac{RTV^2}{P}$$

$$\boxed{V^3 - \left(\frac{RT}{P} + b\right)V^2 + \frac{aV}{P} - \frac{ab}{P} = 0} \rightarrow \text{eqn (1)}$$

⇓

Let 3 roots are V_1, V_2, V_3

at non-critical point

$\boxed{\text{Temp. below } T_c}$

at critical Point, all three Roots are same.

$$\boxed{V_1 = V_2 = V_3 = V_c}$$

$$(V - V_c)^3 = 0$$

$$\boxed{V^3 + 3VV_c^2 - 3V^2V_c + V_c^3 = 0} \rightarrow \text{eqn (2)}$$

compare (1) & (2)

$$\frac{RT_c}{P_c} + b = 3V_c \quad \text{--- (i)}$$

$$\frac{a}{P_c} = 3V_c^2 \quad \text{--- (ii)}$$

$$\frac{ab}{P_c} = V_c^3 \quad \text{--- (iii)}$$

$$\begin{aligned} \text{(ii)} \quad \frac{1}{b} &= \frac{3}{V_c} \\ \text{(iii)} \end{aligned}$$

$$V_c = 3b$$

$$T_c = \frac{8a}{27Rb}$$

$$P_c = \frac{a}{27b^2}$$

Endiometry [Analysis of Gas]

↳ Based on Avagadro's law

↳ at const. T & P $V \propto n$

During Combustion Reaxⁿ (Comp. Reacts with O_2), Gases released are passed through certain solution in which they can be absorbed.

Assumptions used during Endiometric calculations.

- i) If temp. is not given, take $(25^\circ C)$
- ii) Vol. term is used for gases.
- iii) N_2 is inert.
- iv) H_2O vapours can be condensed & Vol. of H_2O vapours ignored.

T_c of all gases have value less than T_c of H_2O .

gas

absorbing Solⁿ

Acidic gases

- $CO_2, SO_2, SO_3, NO_2, X_2$

NaOH or KOH Solⁿ

- CO_2, C_2H_2 etc.

Ammonical Cu_2Cl_2
 $[Cu_2Cl_2 + NH_3(aq)]$

- O_3

Turpentine oil

- NO

$FeSO_4$ solⁿ

- N_2

Heated Mg

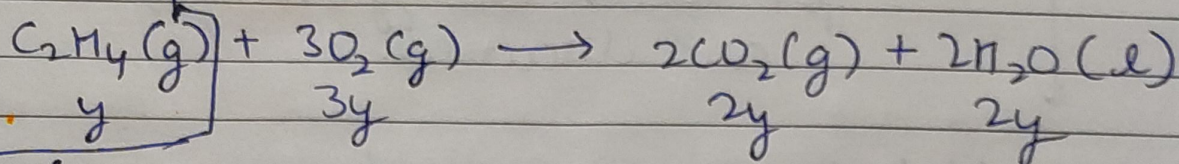
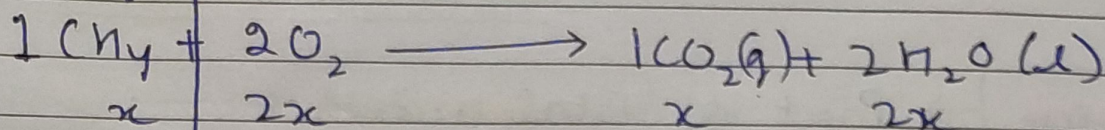
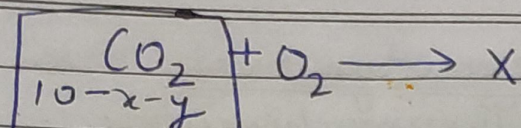
- O_2

Alkaline Pyrogallol

- NH_3
(basic)

Acids

→ 10 ml of a mixture of CH_4, C_2H_4 and CO_2 was exploded with excess of air. After explosion there was a contraction in volume of 17 ml and after treatment with KOH there was further reduction of 14 ml. What was composition of the mixture? (Vol. measurement refers to same P-T cond's).



↓
10ml

$$\begin{aligned} 2x + 2y &= 17 \\ x + 2y &= 14 \\ \hline x &= 3 \end{aligned}$$

$$\begin{aligned} 2x + 2y &= 17 \\ x + 2y &= 14 \\ \hline x &= 3 \end{aligned}$$

$$\boxed{y = 5.5}$$

$$2x + 2y = 17$$

$$10 - x - y + x + 2y = 14$$

$$10 + y = 14$$

$$\boxed{y = 4}$$

$$x = 4.5 \text{ ml}$$

P, T const.

for gases

$V \propto n$

% vol = % mol

Vol. fraction =
mol. fraction

Vol. ratio = mole ratio

- Volume of $\text{CH}_4 \rightarrow 4.5 \text{ ml}$
- Volume of $\text{C}_2\text{H}_4 \rightarrow 4 \text{ ml}$
- Volume of $\text{CO}_2 \rightarrow 1.5 \text{ ml}$